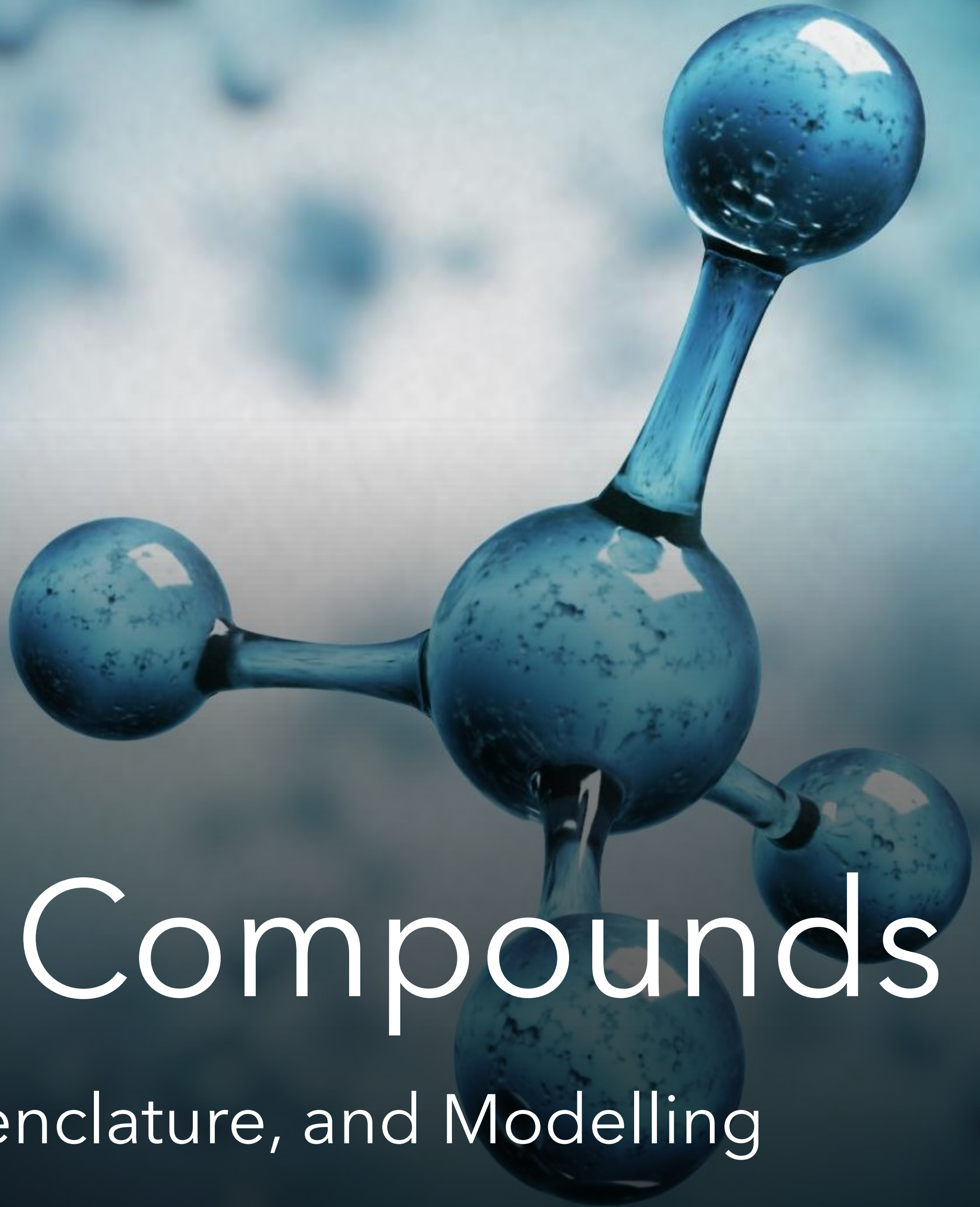


Chemical Compounds

Bond Formation, Nomenclature, and Modelling

Overview

- What are chemical compounds? Why do they form?
- Ionic vs covalent compounds
  - Drawing Bohr models and Lewis diagrams
  - IUPAC naming conventions:
    - Covalent compounds
- Balanced Chemical Equations

Legend (for Sci9PW only):

Do not need to know this slide

Need to know some of what is on this slide; for details, see the "Notes" section of powerpoint.

What are chemical compounds? Why do they form?

Review

1. Why do compounds form?
2. How do you draw the Bohr model for an atom? Ion?
3. What is a valence shell? Valence electron?
4. On the periodic table, where are the metals and non-metals? What is the difference?
5. Which of these compounds are ionic? Covalent? What's the difference?
6. How do you name ionic compounds?

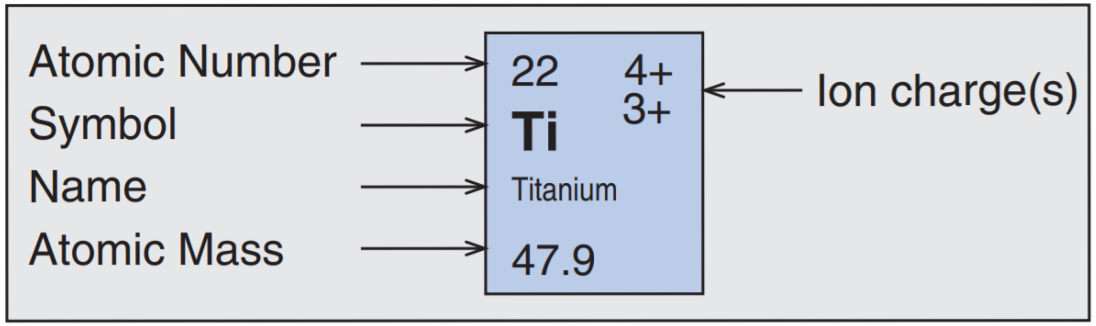
Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the middle of diagram:
   - Element symbol (e.g. "Cl" "F" "Na")
   - # protons, # neutrons
3. Draw the electrons in energy shells:
   - Max electrons per shell from inside to outside: 2, 8, 8, 18
   - Electrons drawn singly starting from top and rotating clockwise
4. Ions only:
   - Add square brackets and a charge

Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.

	protons	**neutrons**	**electrons**
Atom	atomic number	*rounded* atomic mass minus atomic number	atomic number
Ion	atomic number	*rounded* atomic mass minus atomic number	atomic number minus ionic charge

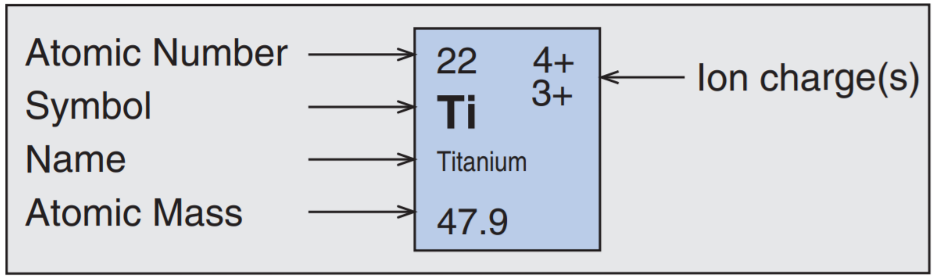


If the tenths place is a 4 or lower, round down.	**32.1 → 32**	**65.4 → 65**
If the tenths place is a 5 or higher, round up.	**10.8 → 11**	**35.5 → 36**

Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.

	protons	neutrons	electrons
Atom	atomic number	*rounded* atomic mass minus atomic number	atomic number
Ion	atomic number	*rounded* atomic mass minus atomic number	atomic number minus ionic charge

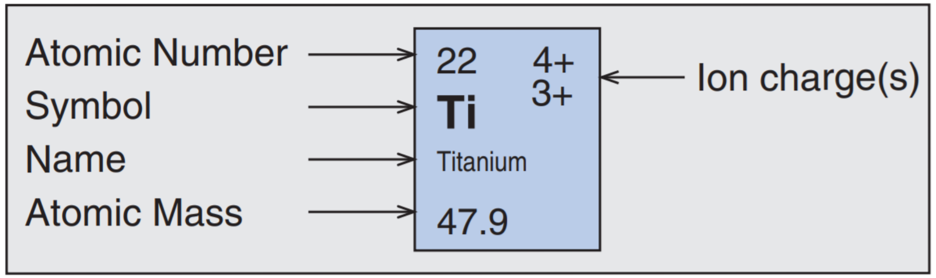


		p	n	e
11 + **Na** Sodium 23.0	Na	11	23-11=12	11
	Na^+	11	23-11=12	11-(+1)=10
12 2+ **Mg** Magnesium 24.3	Mg			
	Mg^{2+}			
8 2– **O** Oxygen 16.0	O			
	O^{2-}			
17 – **Cl** Chlorine 35.5	Cl			
	Cl^-			

Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.

	protons	neutrons	electrons
Atom	atomic number	*rounded* atomic mass minus atomic number	atomic number
Ion	atomic number	*rounded* atomic mass minus atomic number	atomic number minus ionic charge



		p	n	e
11 + Na Sodium 23.0	Na	11	23-11=12	11
	Na^{+}	11	23-11=12	11-(+1)=10
12 2+ Mg Magnesium 24.3	Mg	12	24-12=12	12
	Mg^{2+}	12	24-12=12	12-(+2)=10
8 2– O Oxygen 16.0	O	8	16-8=8	8
	O^{2-}	8	16-8=8	8-(-2)=10
17 – Cl Chlorine 35.5	Cl	17	36-17=19	17
	Cl^{-}	17	36-17=19	18

Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the nucleus:
   - Element symbol
   - # protons, # neutrons
3. Draw the electrons in energy shells:
   - Max electrons per shell from inside to outside: 2, 8, 8, 18
   - (Except in first shell), electrons are filled ***starting at top***, going ***clockwise***, singly at first then paired
4. Ions only:
   - Add square brackets and ion charge from periodic table

Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
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4. Ions only:
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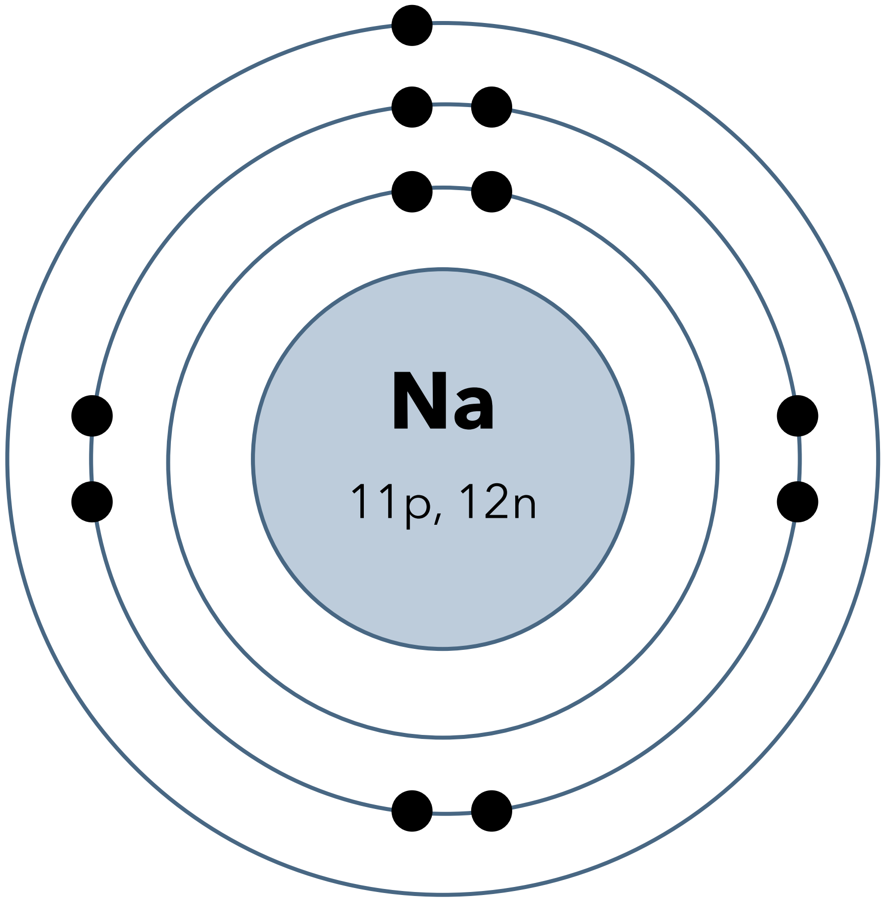
	p	n	e
Na	11	23-11=12	11

Example: sodium atom



Review: Drawing Bohr Models of Atoms and Ions

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4. Ions only:
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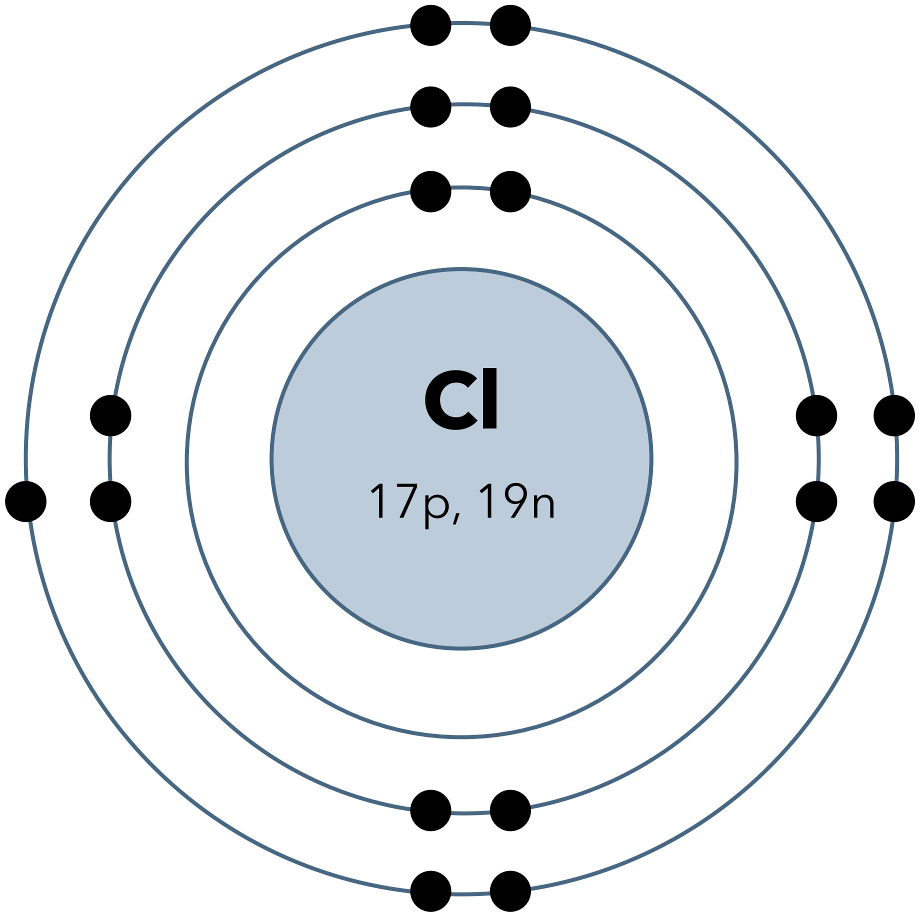
	p	n	e
Cl	17	36-17=19	17

Example: chlorine atom



Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the nucleus:
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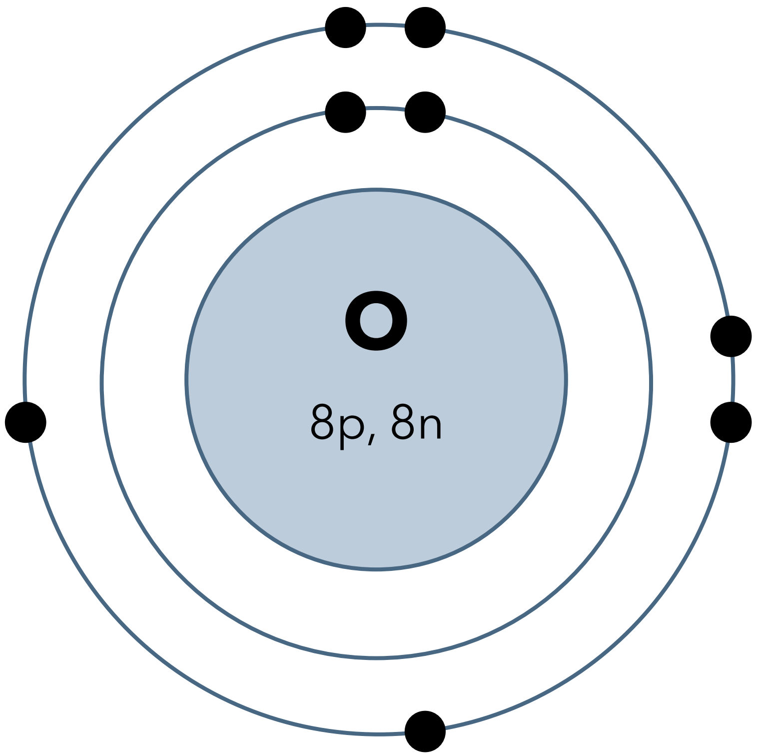
   - # protons, # neutrons
3. Draw the electrons in energy shells:
   - Max electrons per shell from inside to outside: 2, 8, 8, 18
   - (Except in first shell), electrons are filled starting at top, going clockwise, singly at first then paired
4. Ions only:
   - Add square brackets and ion charge from periodic table

	p	n	e
O	8	16-8=8	8

Example: oxygen atom



Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the nucleus:
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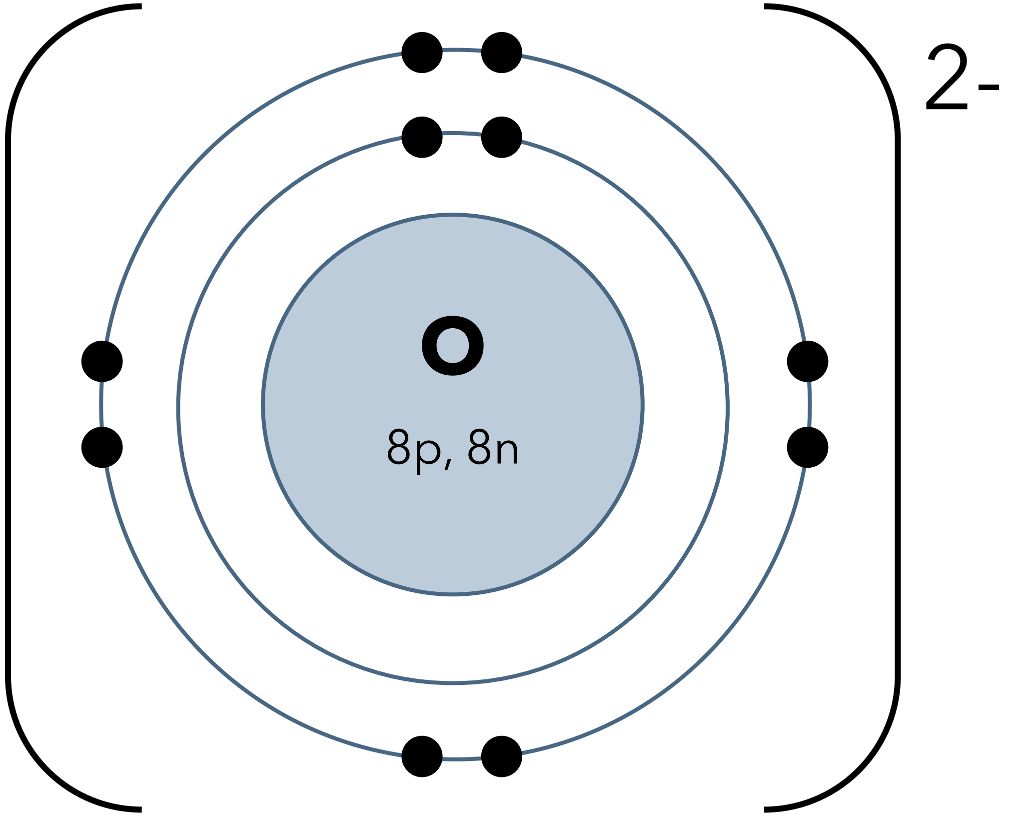
4. Ions only:
   - Add square brackets and ion charge from periodic table

	p	n	e
O^{2-}	8	16-8=8	8-(-2)=10

Note: subtracting a negative is the same as adding.

Example: oxygen ion



Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the nucleus:
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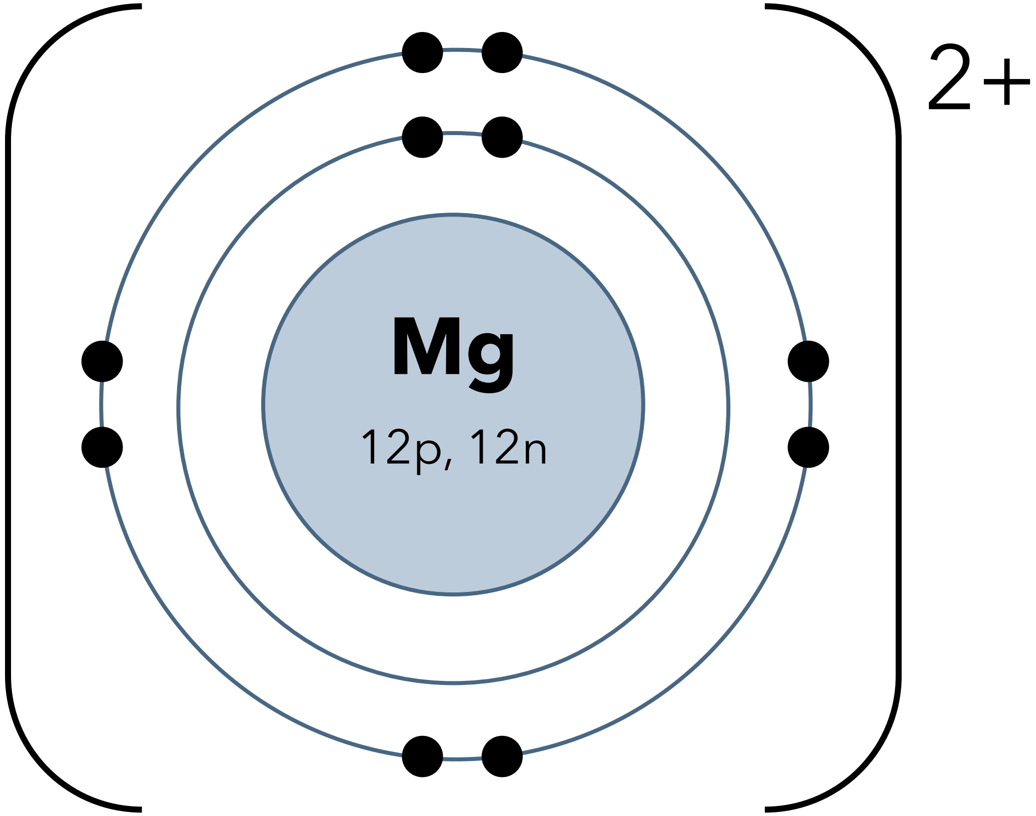
4. Ions only:
   - Add square brackets and ion charge from periodic table

	p	n	e
Mg^{2+}	12	24-12=12	12-(+2)=10

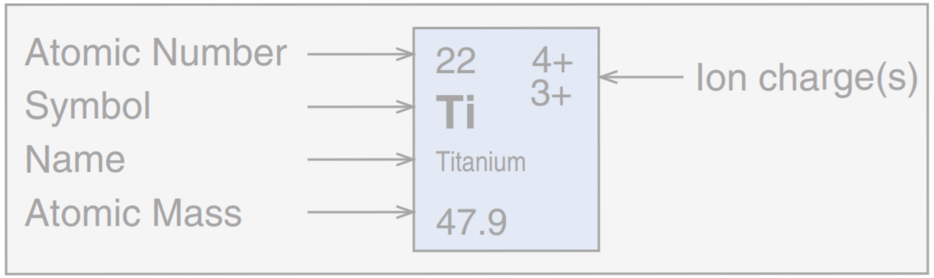
Example: magnesium ion



Review: Drawing Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.

	protons	neutrons	electrons
Atom	atomic number	atomic number minus *rounded* atomic mass	atomic number
Ion	atomic number	atomic number minus *rounded* atomic mass	atomic number minus ionic charge



		p	n	e
11 + Na Sodium 23.0	Na	11	23-11=12	11
	Na^{+}	11	23-11=12	11-(+1)=$\boxed{10}$
12 2+ Mg Magnesium 24.3	Mg	12	24-12=12	12
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	O^{2-}	8	16-8=8	8-(-2)=$\boxed{10}$
17 –	Cl	17	36-17=19	17

How come so many of the ions have the same number of electrons? What is an ion, anyways?

Achieving Stability Through Nobility

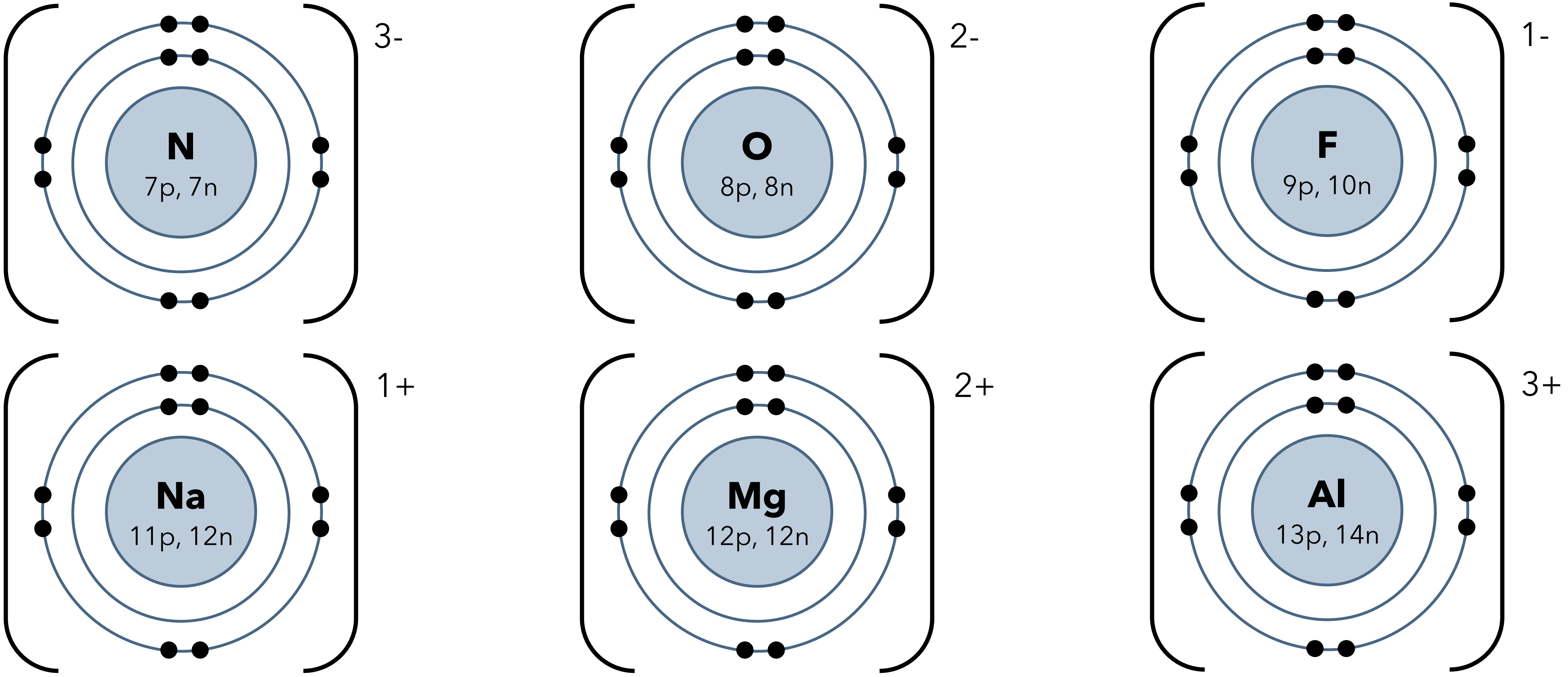
Activity:

1. Draw the Bohr model for one of the following ions.

N^{3-}	O^{2-}	F^{-}
Na^{+}	Mg^{2+}	Al^{3+}

2. Compare your Bohr model with other students in the class. What do they have in common? What is different?

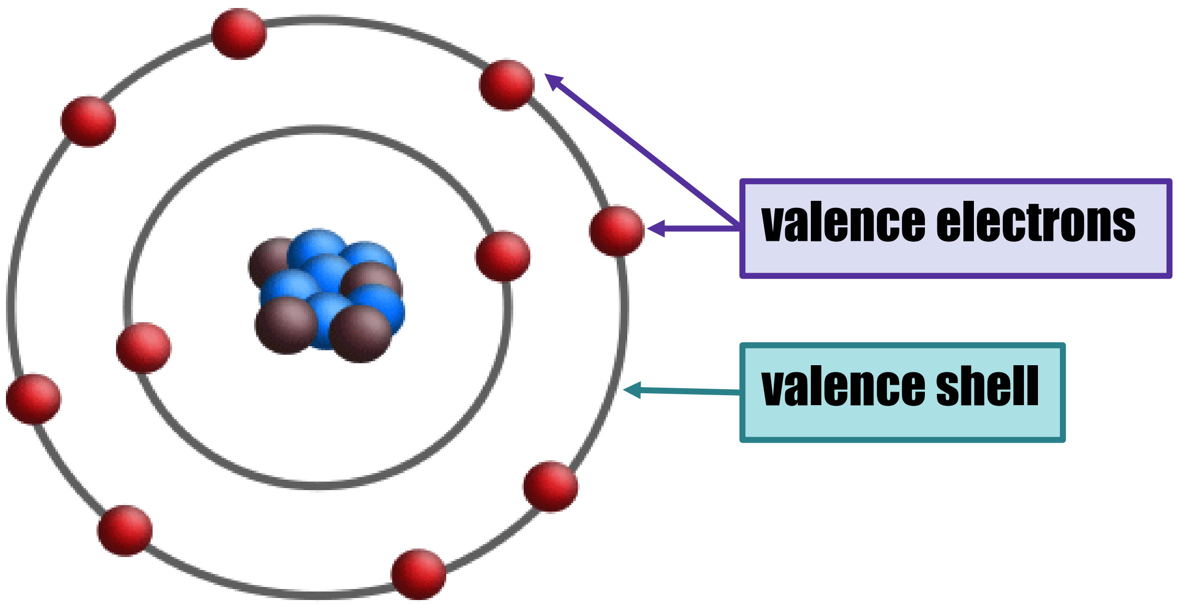
Achieving Stability Through Nobility



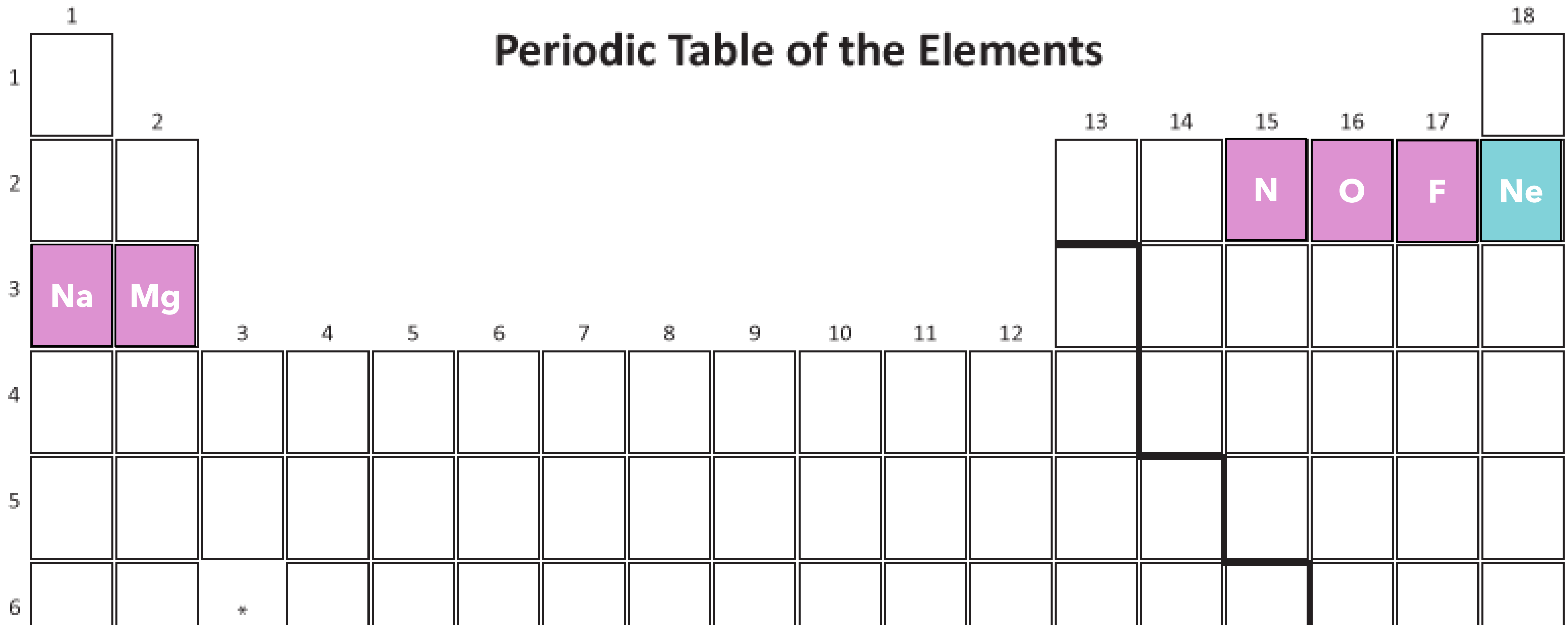
2. Compare the Bohr models. What do they have in common? What is different?

Achieving Stability Through Nobility

- The **valence shell** is the outermost shell containing electrons. Electrons in this shell are called **valence electrons**.
- A stable atom has a full valence shell.



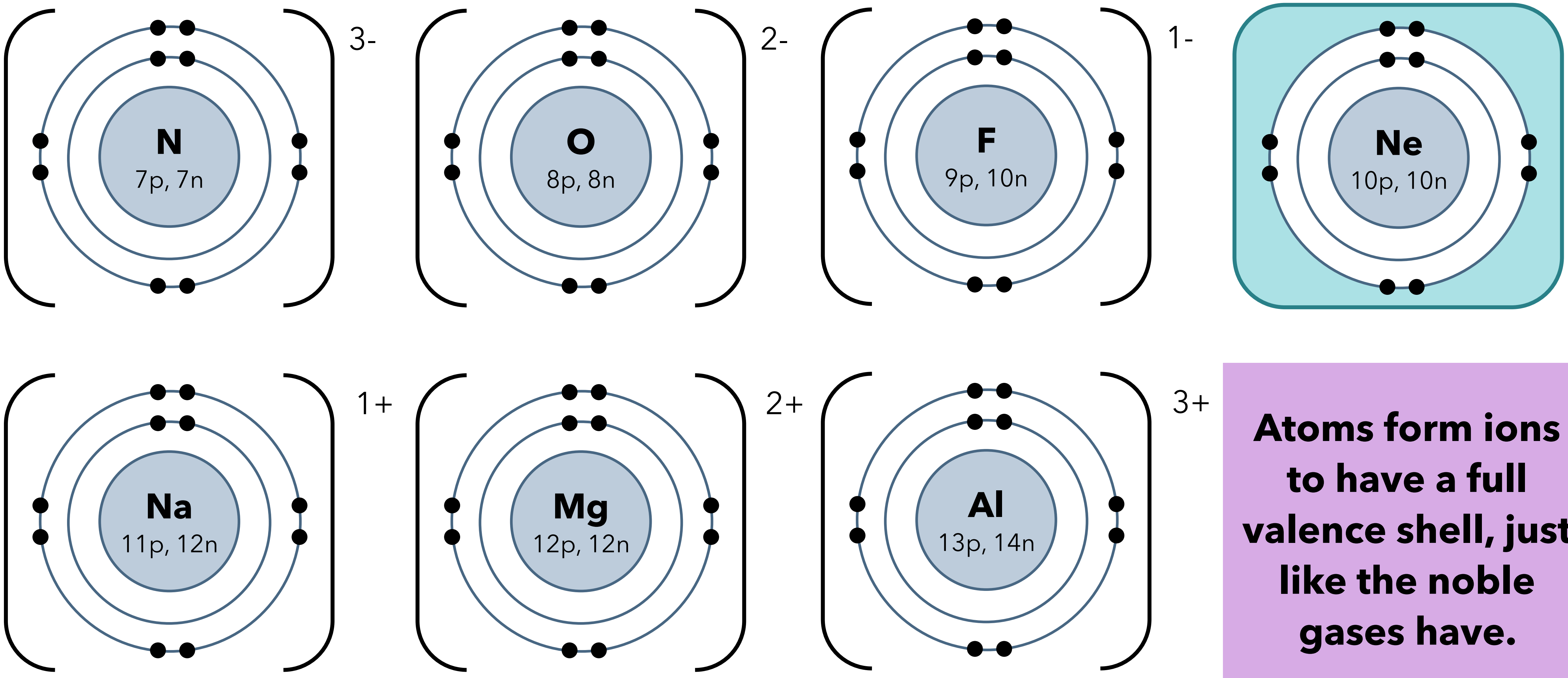
Achieving Stability Through Nobility



Achieving Stability Through Nobility

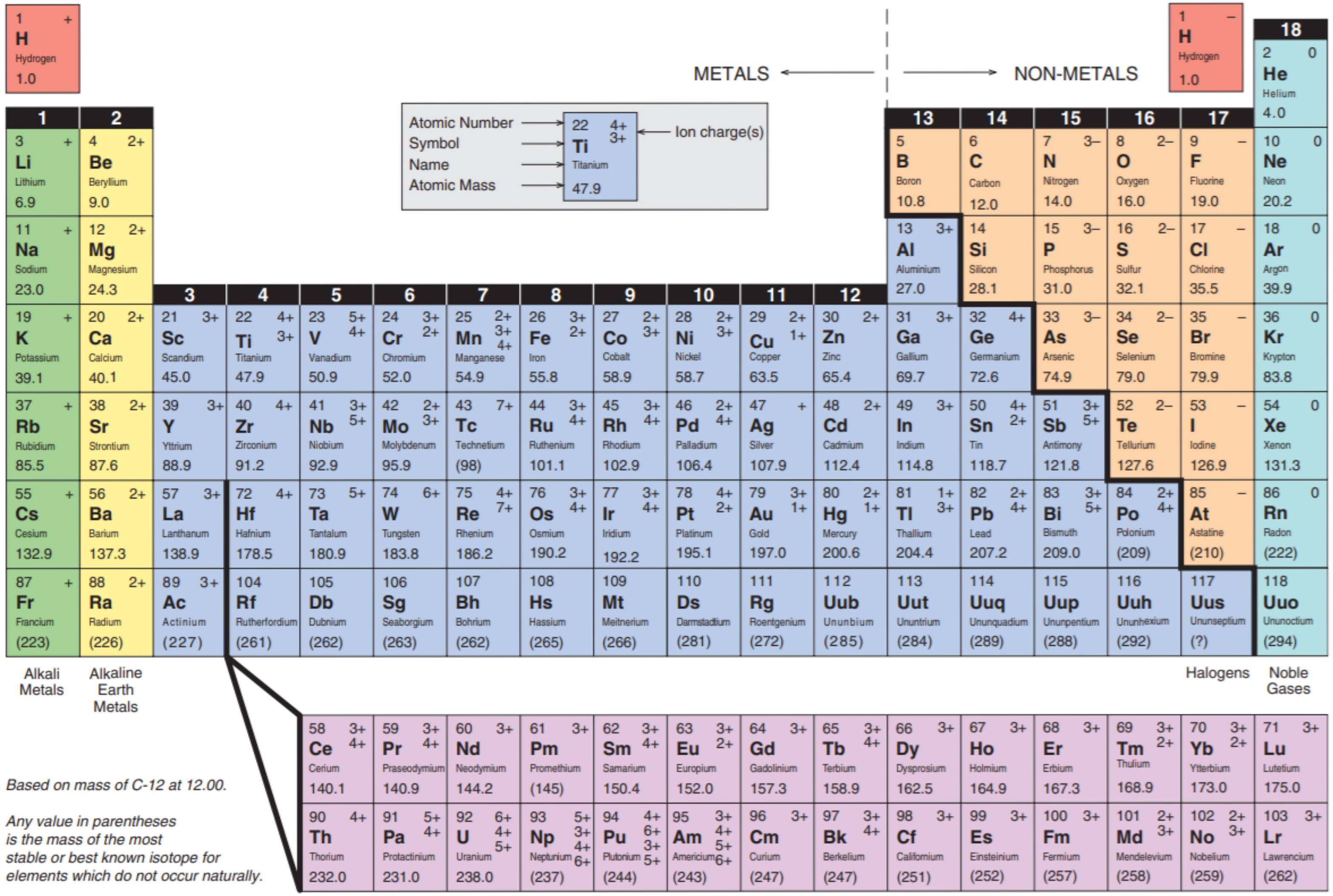
- Atoms form **compounds** to have a full valence shell.
  - **Ionic compound**: atoms ***gain or lose electrons***
  - **Covalent compound**: atoms ***share electrons***

Achieving Stability Through Nobility



Atoms form ions to have a full valence shell, just like the noble gases have.

Achieving Stability Through Nobility



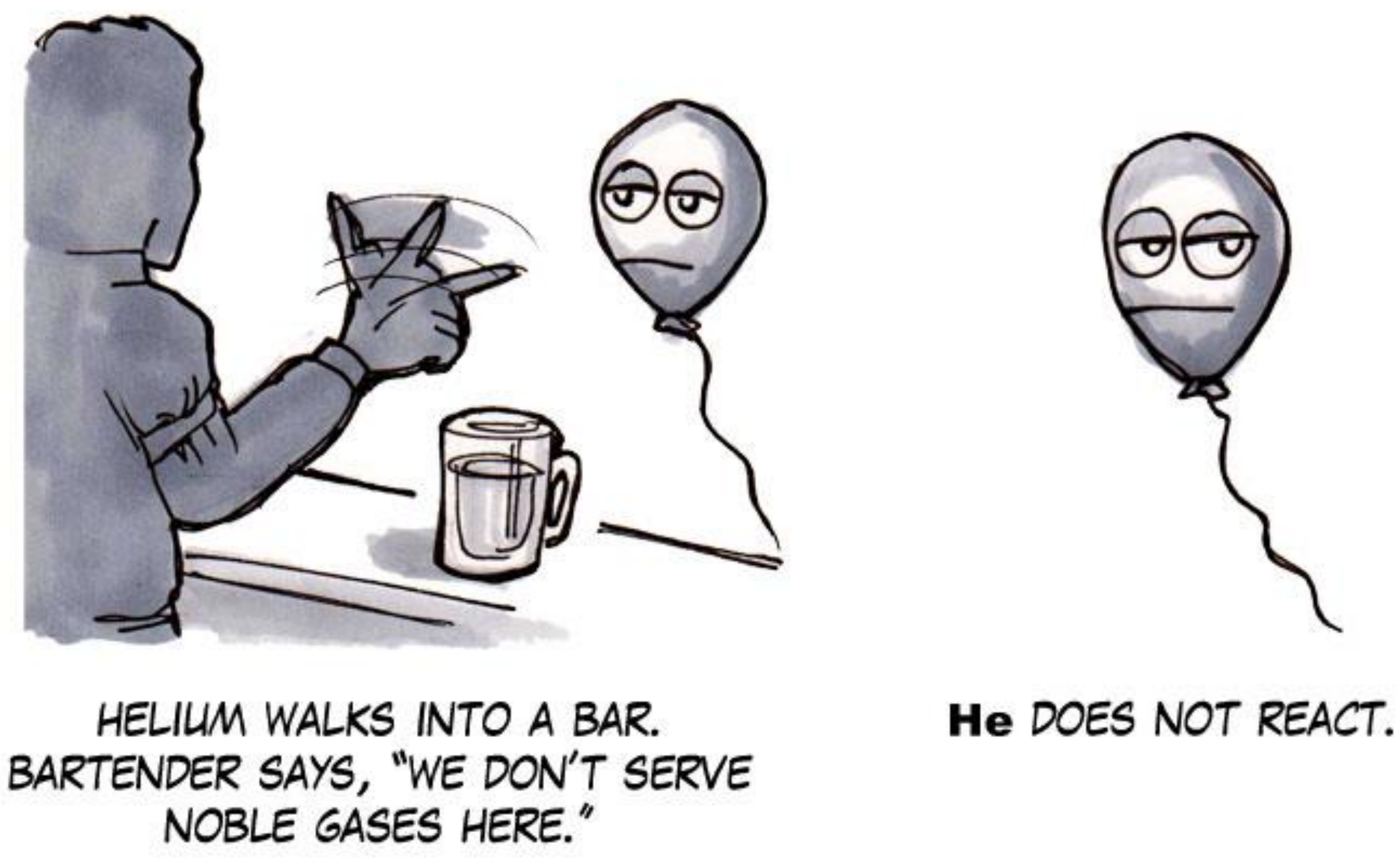
Valence shells can also be used to explain reactivity.

Alkali metals and halogens extremely reactive: only 1 electron away from full valence shell.

Alkaline earth metals and Group 16 elements very reactive: 2 electrons away.

Noble gases non-reactive.

Achieving Stability Through Nobility



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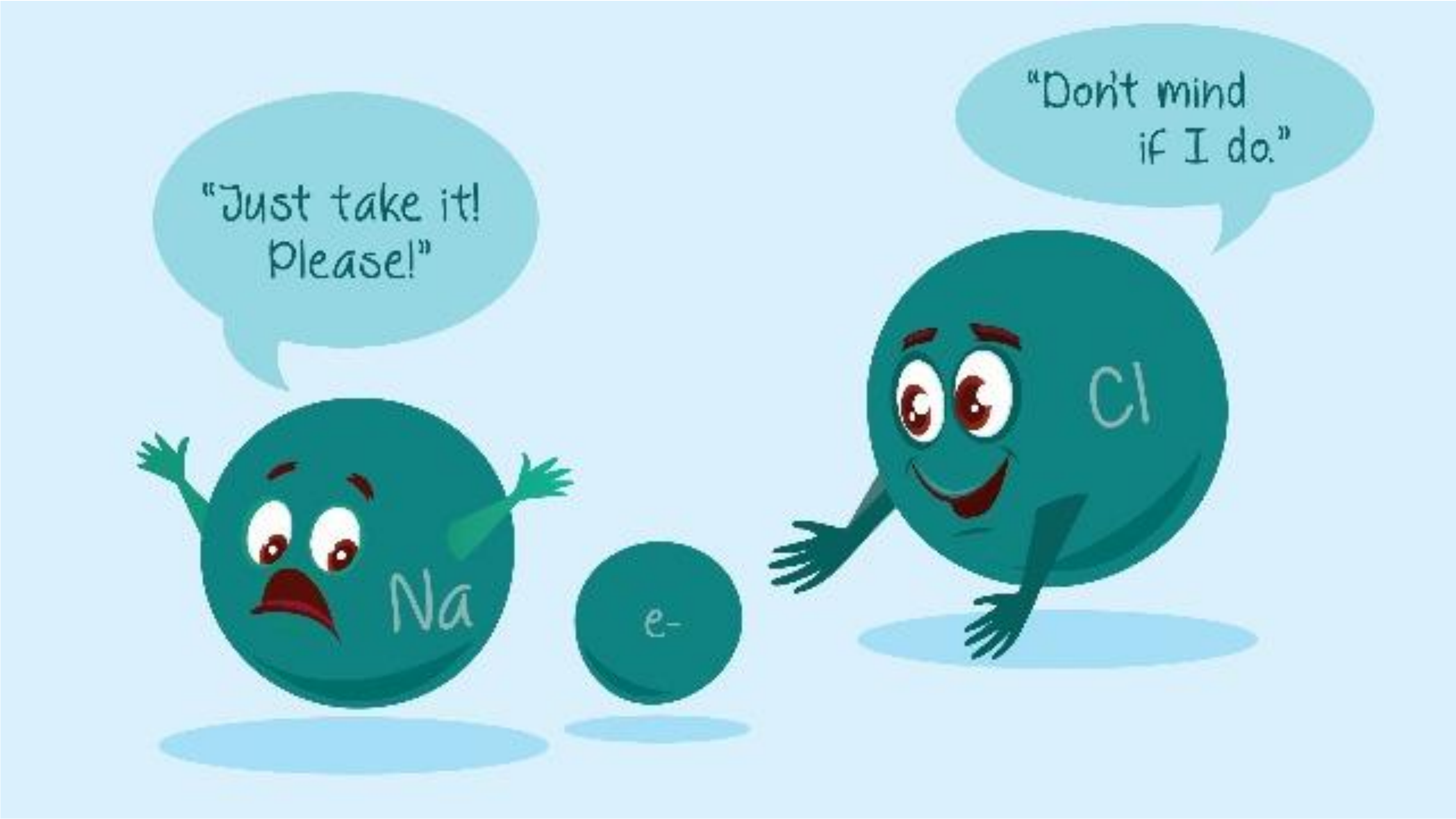
Ionic Compound Formation



https://www.technologynetworks.com/analysis/articles/cation-vs-anion-definition-chart-and-the-periodic-table-322863

Ionic Compound Formation

- Atoms form ions to have a **full valence shell**, just like the noble gases have.
- Electrons are negatively charged. When electrons are added or taken away, atoms become positively or negatively charged ions.
  - **Cation**: positively charged ion (e.g. Ca^{2+}, Cr^{3+}, NH_4^+); forms when *electrons are lost* from an atom
  - **Anion**: negatively charged ion (e.g. N^{3-}, S^{2-}, PO_4^{3-}); forms when *electrons are gained* by an atom

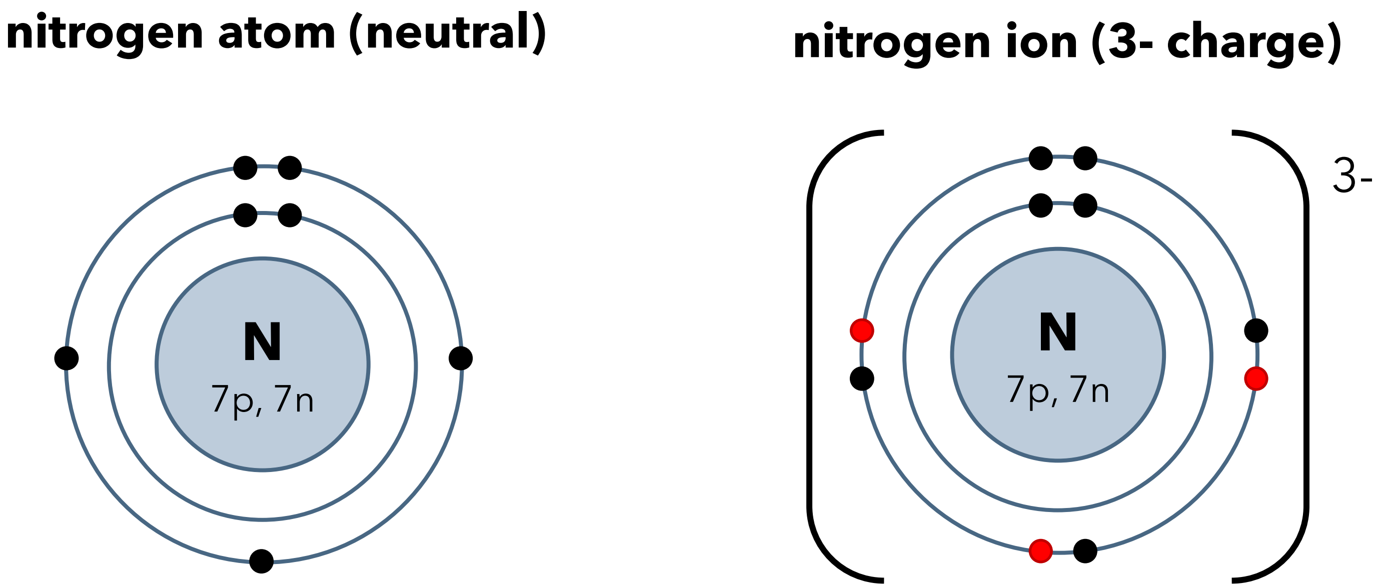
Note: NH_4^+ and PO_4^{3-} are **polyatomic ions** because they consist of multiple ("*poly-*") atoms ("*-atomic*").

Ionic Compound Formation

https://sciencescribbles.tumblr.com/post/136488755835/chapter-21-the-nuclear-atom-part-2-variations

Ionic Compound Formation

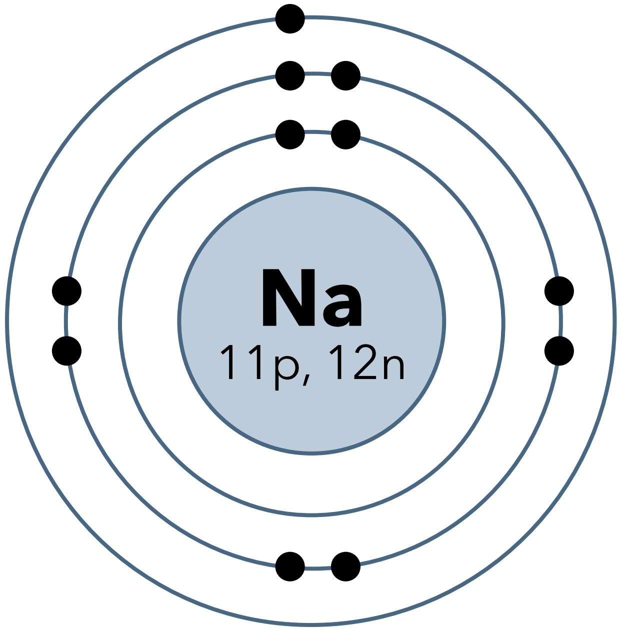
- Atoms are neutral because #protons = #electrons.
- Nitrogen atom becomes an ion when it gains 3 electrons.



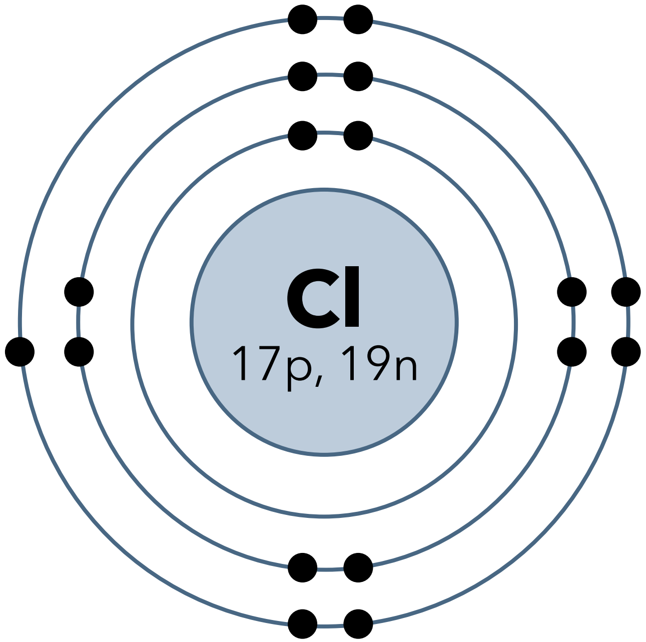
Where do these electrons come from?

Ionic Compound Formation (NaCl)

- Ionic compounds form when **electrons are transferred** and ions are formed. Usually involves a **metal and a non-metal**.



sodium atom (neutral)



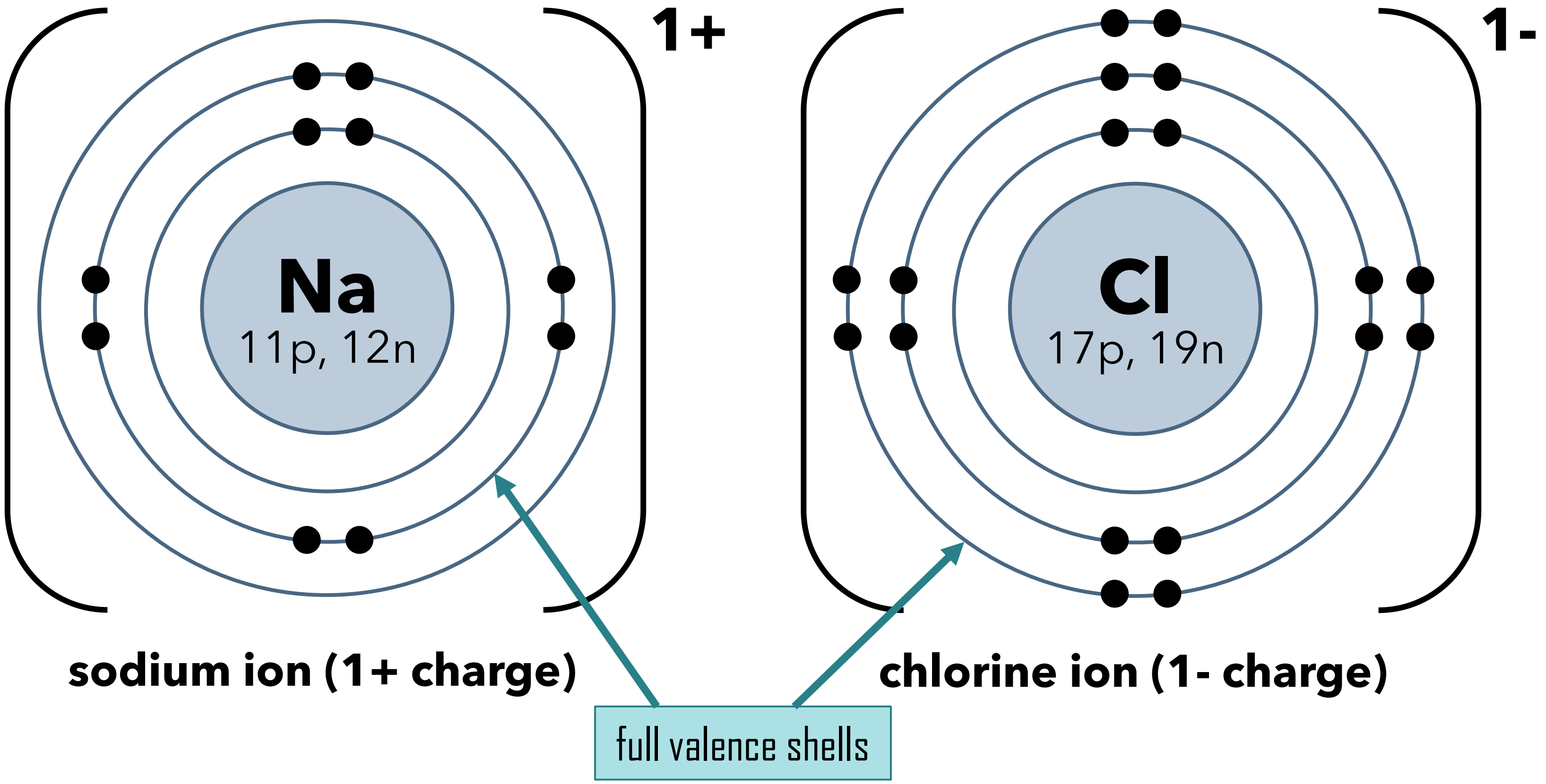
chlorine atom (neutral)

In order to get full valence shells:

- Na needs to **lose 1** electron.
- Cl needs to **gain 1** electron.

Ionic Compound Formation (NaCl)

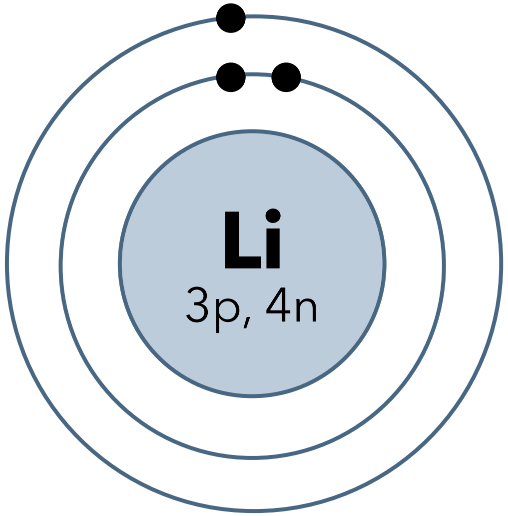
- Ionic compounds form when **electrons are transferred** and ions are formed. Usually involves a **metal and a non-metal.**



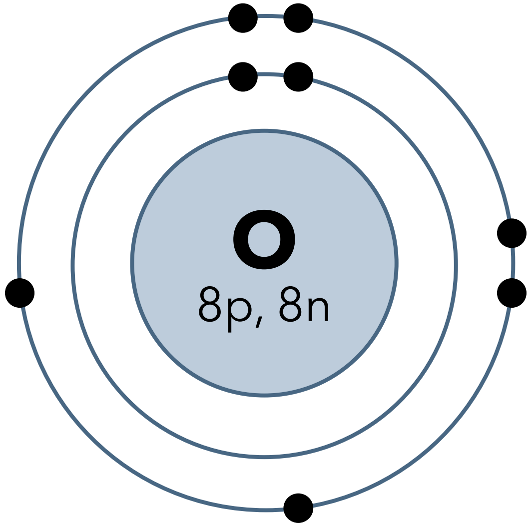
This ionic compound is **NaCl** (sodium chloride). It has one Na^+ ion and one Cl^- ion.

Ionic Compound Formation (Li_2O)

- Ionic compounds form when **electrons are transferred** and ions are formed. Usually involves a **metal and a non-metal**.



lithium atom (neutral)



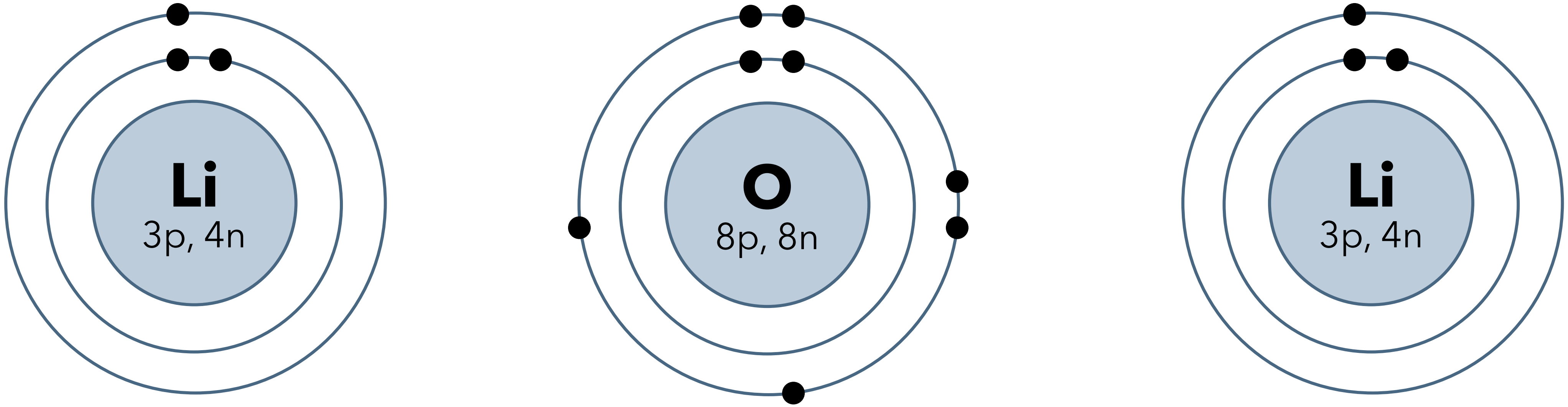
oxygen atom (neutral)

- Li needs to **lose 1** electron.
- O needs to **gain 2** electrons.

Problem: Electron numbers not balanced.

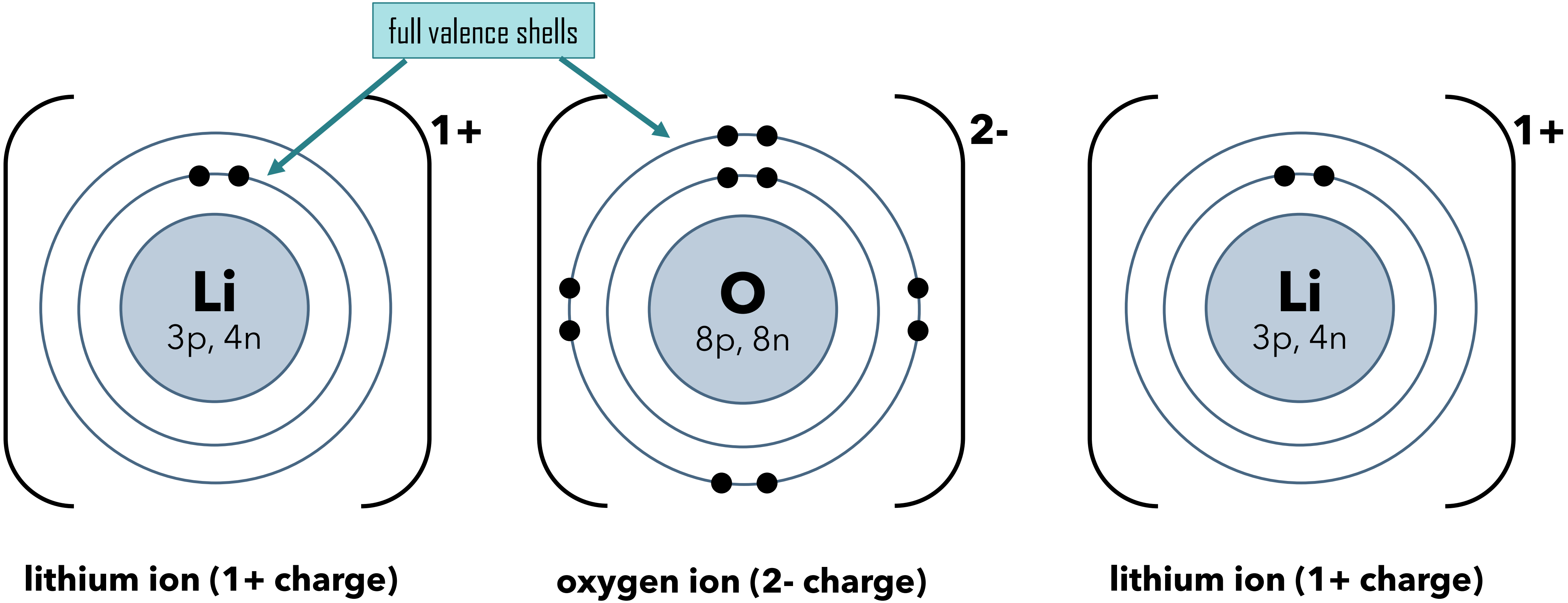
Solution: The compound needs two lithium ions!

Ionic Compound Formation (Li_2O)



lithium atom (neutral) **oxygen atom (neutral)** **lithium atom (neutral)**

Ionic Compound Formation (Li_2O)



This ionic compound is **Li_2O** (lithium oxide). It has two Li^+ ions and one O^{2-} ion.

Subscripts

Subscripts in Chemical Compounds

- **Subscripts** are small numbers written on the bottom right of an element or ion to show how many are in that compound.
- **No subscript** means there is *only one* of that element or ion.
- A subscript outside a bracket indicates multiples of a polyatomic ion (*multiply* subscripts!).

$$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$$

subscripts

Subscripts in Chemical Compounds

Important! If there is **no subscript**, this means there is only **one** of that element or ion.

Chemical Formula	How Many Atoms?
N_2O_3	2 nitrogen 3 oxygen
PF_4	1 phosphorus 4 fluorine
Li_2O	2 lithium ions 1 oxygen ion
Ni_2S_3	2 nickel ions 3 sulfur ions

Subscripts in Chemical Compounds

Practice!

Chemical Formula	How Many Atoms?
Co_2S_3	
PF_4	
$MgBr_2$	
Be_3N_2	

Chemical Formula	How Many Atoms?
H_2O	
CCl_4	
$CaCO_3$	
NaOH	

Bohr Models of Ionic Compounds

Bohr Models of Ionic Compounds

1. Determine how many of each ion is in the compound, from the subscripts.
2. Use the periodic table to find the ionic charge of each ion.
3. Draw the Bohr models of all the ions in the compound. (They should all have full valence shells.)

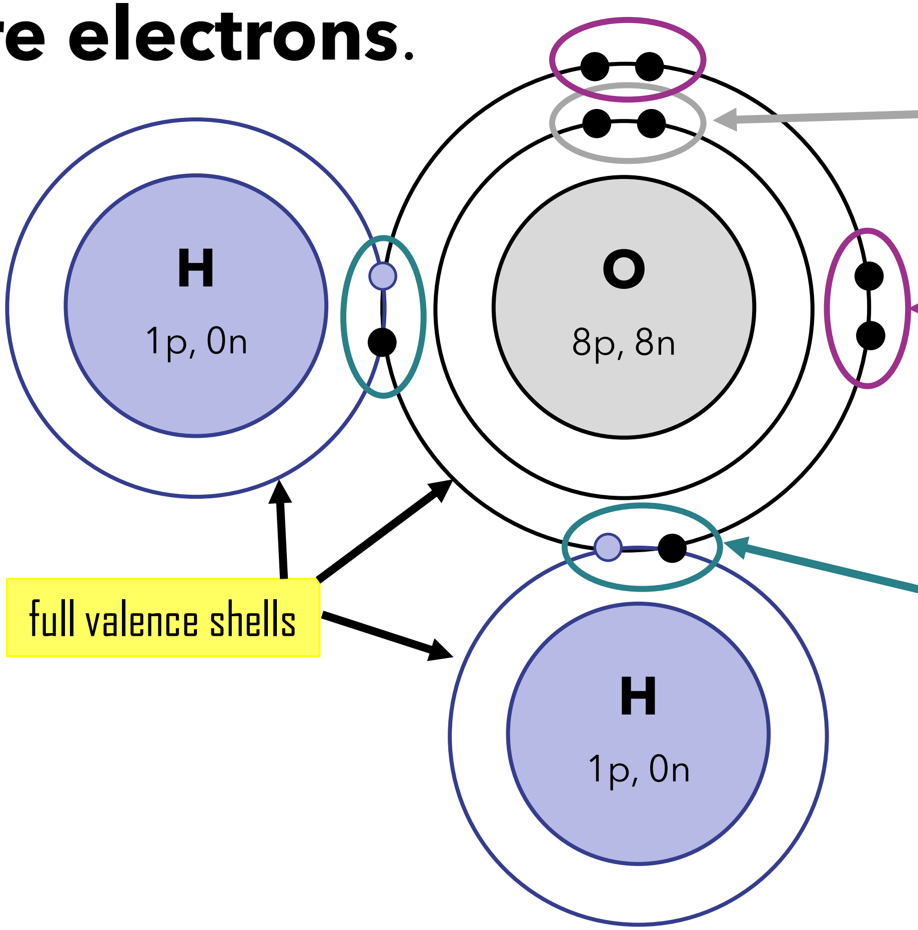
Practice:

a) $MgCl_2$
b) Li_3N

Covalent Compound Formation

- Covalent compounds form when two (or more) **non-metal** atoms **share electrons**.

This covalent compound is **H_2O** (water or dihydrogen monoxide). It has two hydrogen atoms and one oxygen atom.



(These electrons are not in the valence shell. This is *not* a lone pair.)

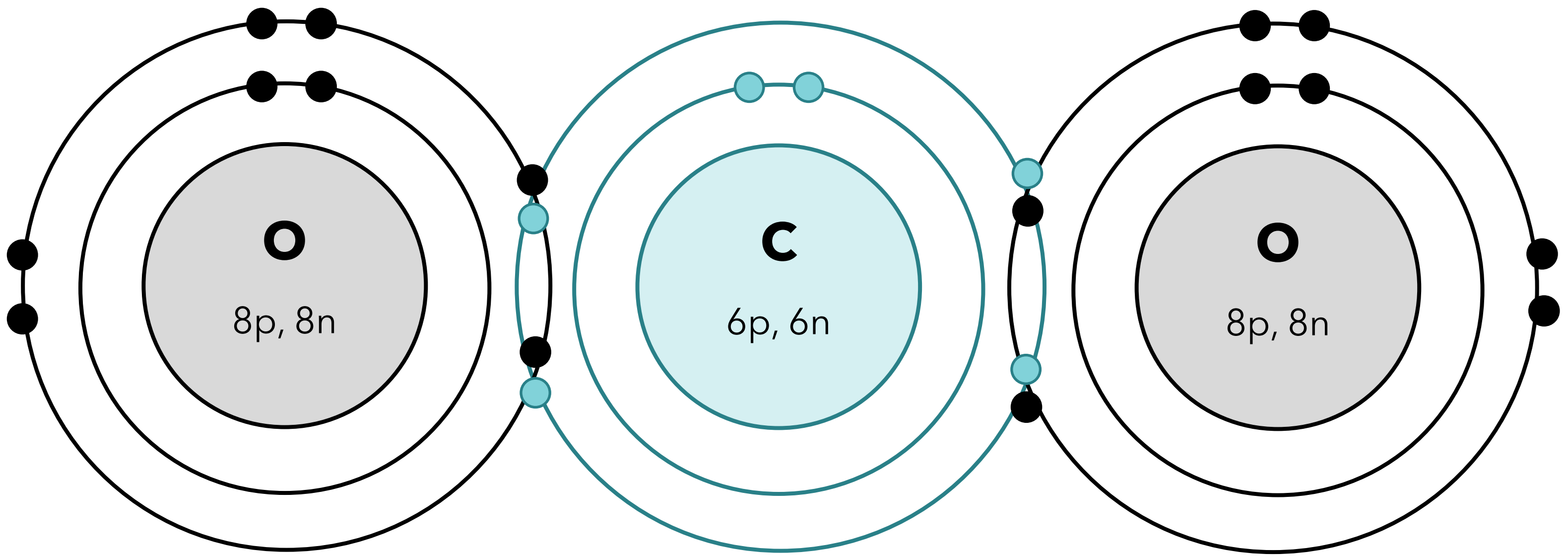
Lone pair: pair of valence electrons that is not shared between atoms

Bonding pair: shared pair of valence electrons in a covalent compound

Total: 2 bonding pairs, 2 lone pairs

Covalent Compound Formation

- Covalent compounds form when two (or more) **non-metal** atoms **share electrons**.



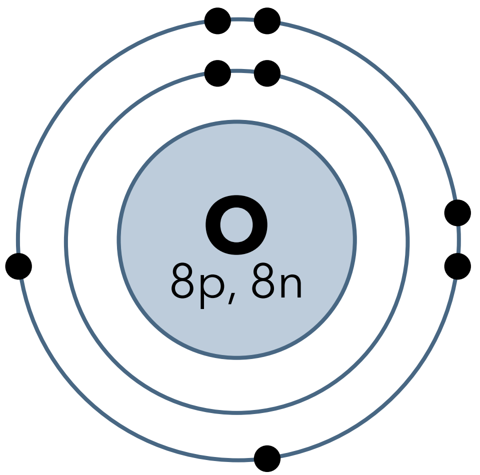
This covalent compound is **CO_2** (carbon dioxide).
It has one carbon atom and two oxygen atoms.

Total: 4 bonding pairs, 4 lone pairs

Introducing Lewis Structures

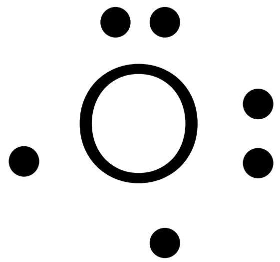
Bohr Model

- All electrons
- All energy shells
- Shows protons and neutrons
- Shows a lot of information, but is clunky and time-consuming

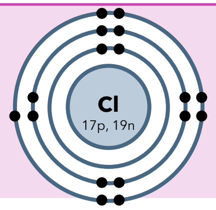
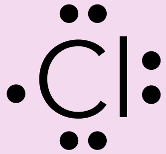
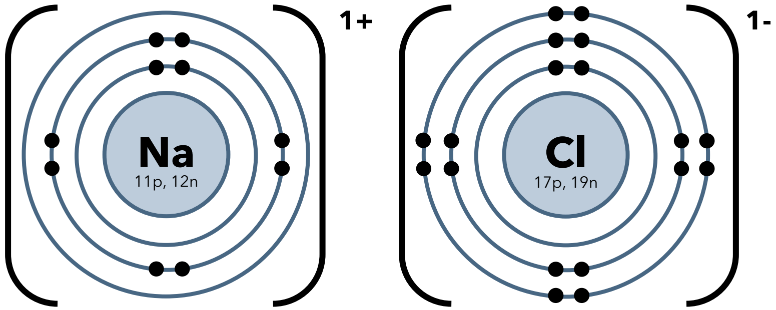
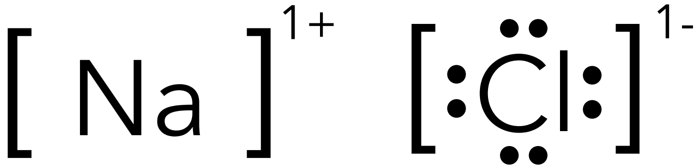
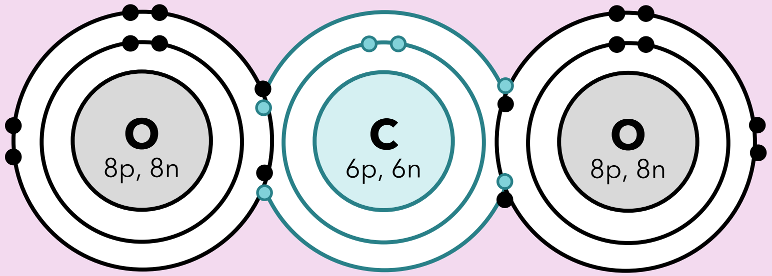
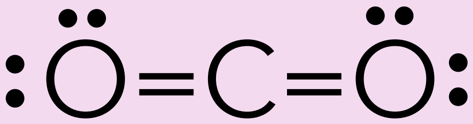


Lewis Structure

- Only valence electrons (except cations)
- Outermost shell only
- Protons and neutrons ignored
- Good at determining bonding in a covalent compound



Introducing Lewis Structures

	Bohr Model	Lewis Structure
Atom		
Ionic Compound		*(not testable)* 
Covalent Compound		

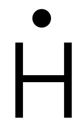
Lewis Structures of Atoms

1. Write element symbol (capitalization matters!)
2. Draw valence electrons around, using the same positions as the Bohr model (i.e. clockwise, unpaired at first then paired)

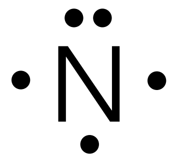
Practice: Draw the Lewis structures of:

a) Mg atom

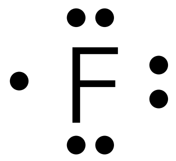
c) H atom



b) N atom

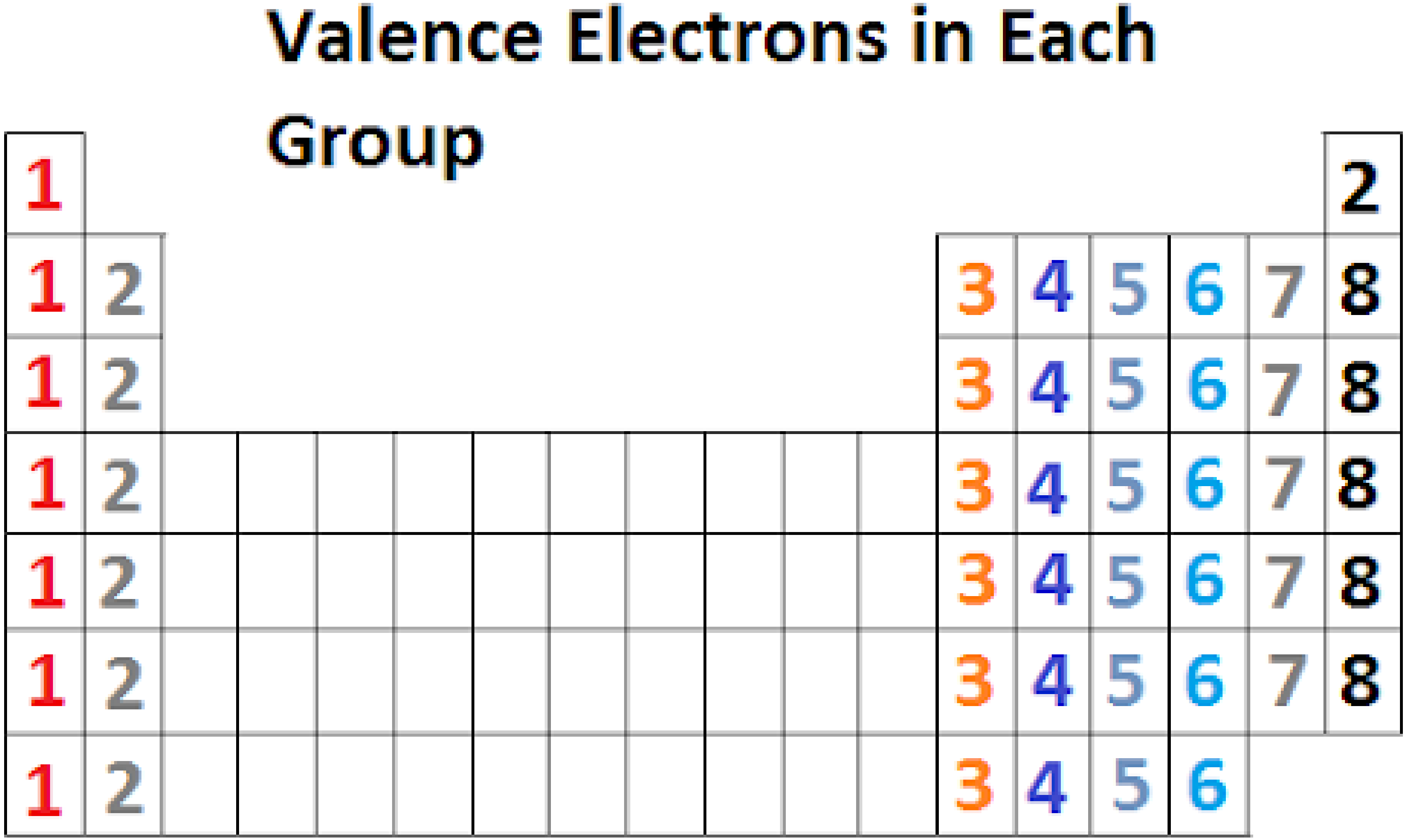


d) F atom



What is a fast way to figure out the number of valence electrons in an atom?

Lewis Structures of Atoms



Look at the last digit of the group #. Exception: hydrogen and helium.

Lewis Structures of Covalent Compounds

Rule 1: All electrons (from the bonded atoms) must be used.

Rule 2: All atoms must have a full valence shell.

1. Draw the Lewis structure of each atom. (Count how many electrons you have in total; write this down.)
2. Determine how many bonds each atom "needs" to complete its valence shell.
3. Guess and check with single, double, and triple bonds until your structure satisfies Rule 1 AND Rule 2.

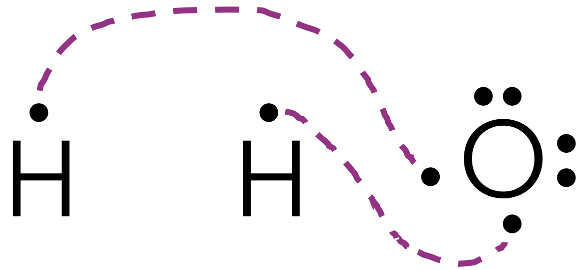
Lewis Structures of Covalent Compounds

Rule 1: All electrons (from the bonded atoms) must be used.

Rule 2: All atoms must have a full valence shell.

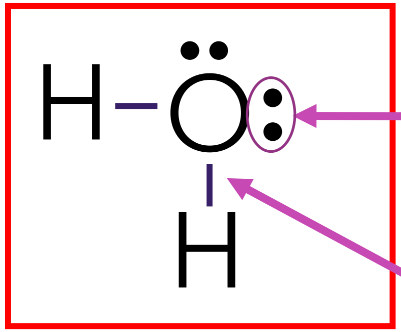
Example: H_2O

1. Draw the Lewis structure of each atom. (Count how many electrons you have in total; write this down.)
2. Determine how many bonds each atom "needs" to complete its valence shell.
3. Guess and check with single, double, and triple bonds until your structure satisfies Rule 1 AND Rule 2.



Each H needs 1 bond;
O needs 2 bonds.

Total e = 8



This is a lone pair.

This is a single bond. It represents a bonding pair of electrons.

2 lone pairs;
2 bonding pairs

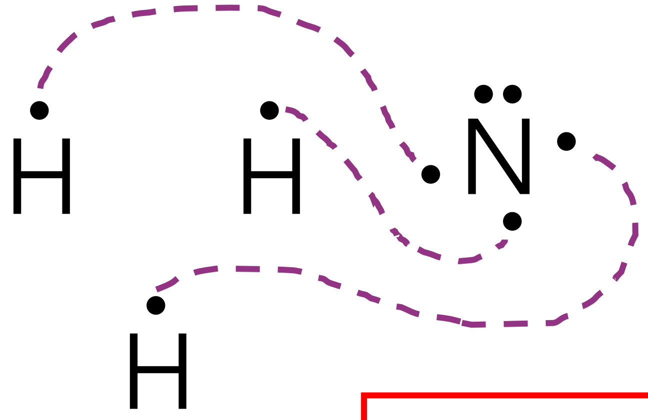
Lewis Structures of Covalent Compounds

Rule 1: All electrons (from the bonded atoms) must be used.

Rule 2: All atoms must have a full valence shell.

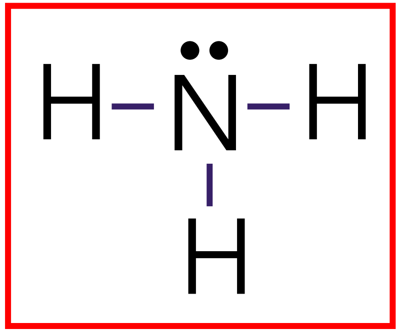
Example: NH_3

1. Draw the Lewis structure of each atom. (Count how many electrons you have in total; write this down.)
2. Determine how many bonds each atom "needs" to complete its valence shell.
3. Guess and check with single, double, and triple bonds until your structure satisfies Rule 1 AND Rule 2.



Each H needs 1 bond; N needs 3 bonds.

Total e = 8



1 lone pair;
3 bonding pairs

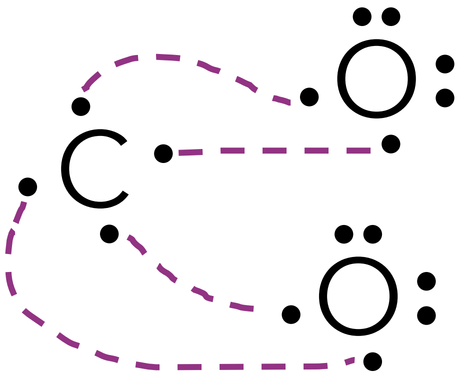
Lewis Structures of Covalent Compounds

Rule 1: All electrons (from the bonded atoms) must be used.

Rule 2: All atoms must have a full valence shell.

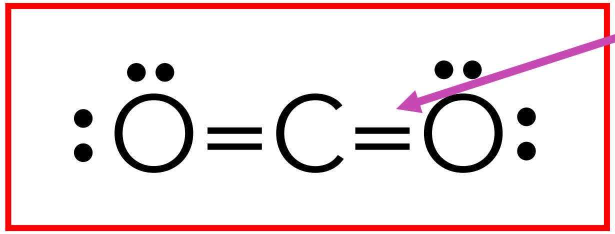
Example: CO_2

1. Draw the Lewis structure of each atom. (Count how many electrons you have in total; write this down.)
2. Determine how many bonds each atom "needs" to complete its valence shell.
3. Guess and check with single, double, and triple bonds until your structure satisfies Rule 1 AND Rule 2.



C needs 4 bonds; each O needs 2 bonds.

Total e = 16



This is a double bond. It represents two bonding pairs of electrons.

4 lone pairs;
4 bonding pairs

Lewis Structures of Covalent Compounds

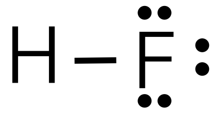
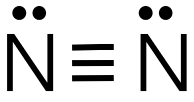
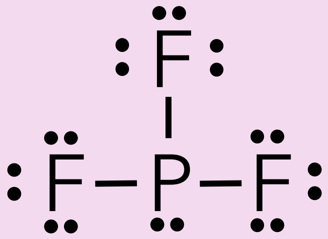
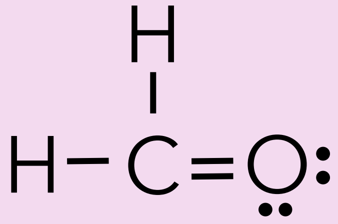
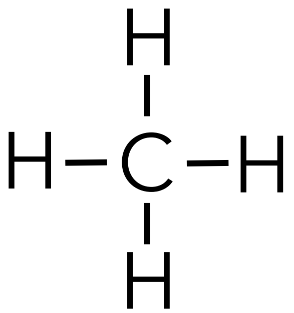
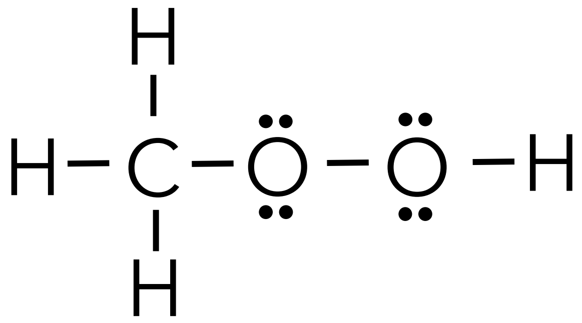
Try drawing the following covalent compounds!

- HF
- PF_3
- CH_4
- N_2 *
- CH_2O
- CO_2H_4 *(challenge)*

*Technically, N_2 is not a compound because it is only made of one element. But, the bonds between the atoms are covalent so we can still draw its Lewis structure.

Lewis Structures of Covalent Compounds

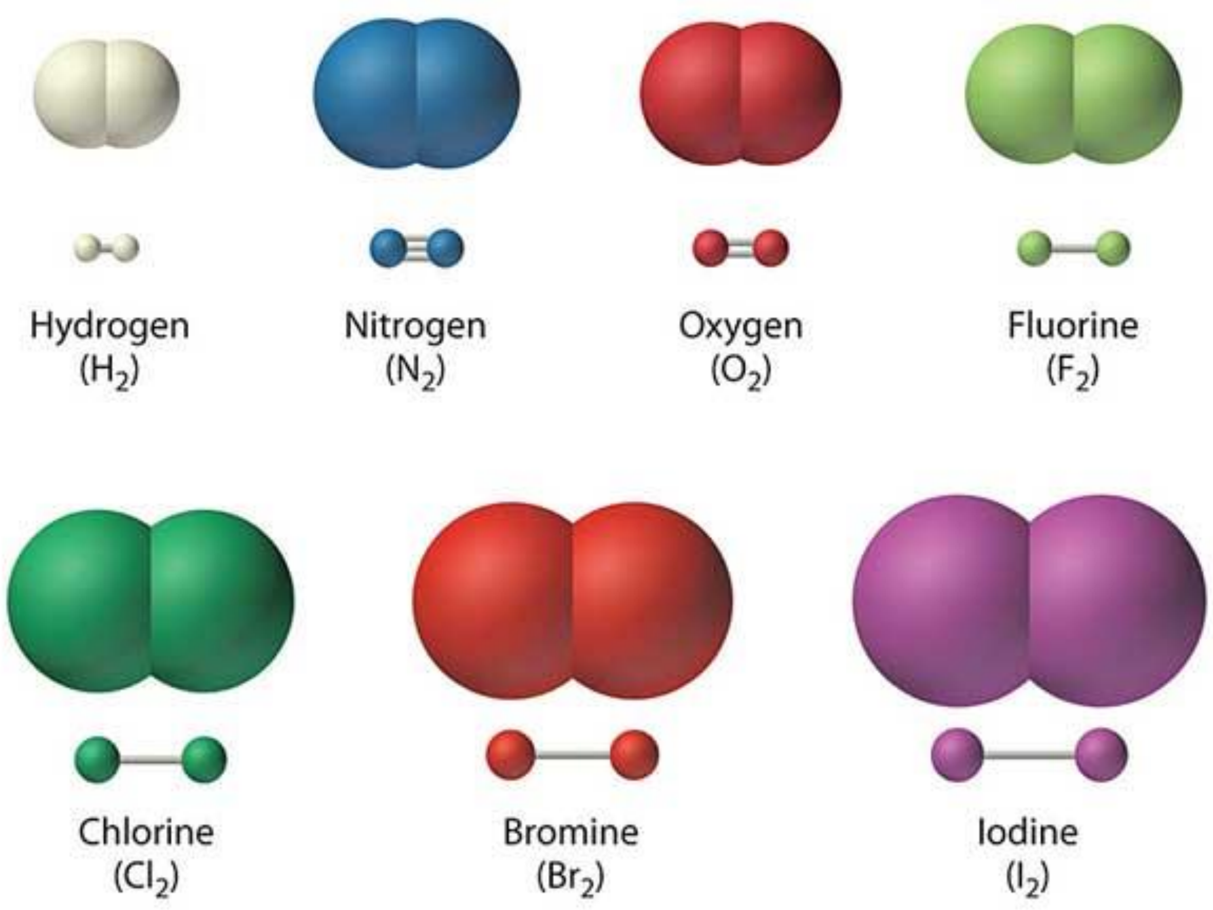
Try drawing the following covalent compounds!

	**HF** (3 lone pairs; 1 bonding pair)		**N_2** * (2 lone pairs; 3 bonding pairs)
	**PF_3** (10 lone pairs; 3 bonding pairs)		**CH_2O** (2 lone pairs; 4 bonding pairs)
	**CH_4** (0 lone pairs; 4 bonding pairs)		**CO_2H_4 (challenge)** (4 lone pairs; 6 bonding pairs)

*Technically, N_2 is not a compound because it is only made of one element. But, the bonds between the atoms are covalent so we can still draw its Lewis structure.

Revisiting Diatomic Elements

- When in their elemental (i.e. not in a compound) form, these elements exist as **diatomic molecules**: two atoms bonding covalently to fill their valence shells.
- Must memorize!



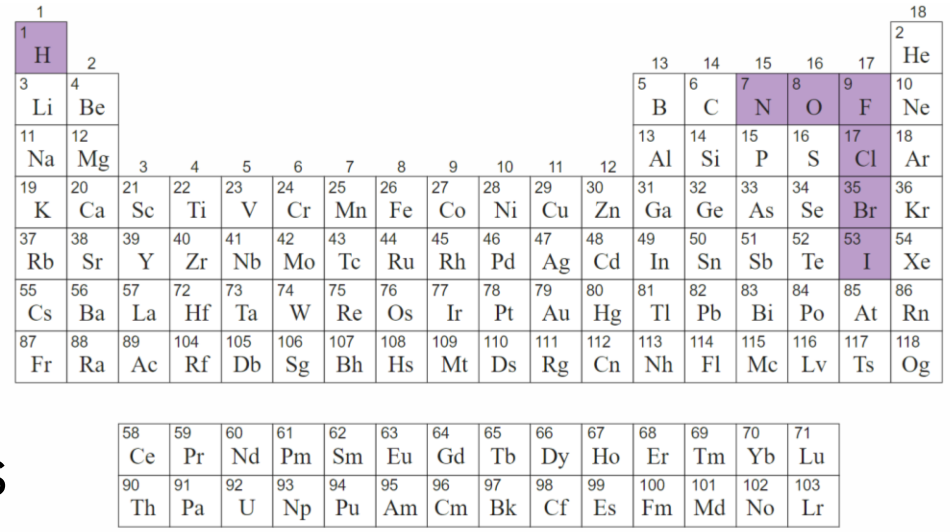
https://fl-pda.org/independent/courses/elementary/science/section4/4c6.htm

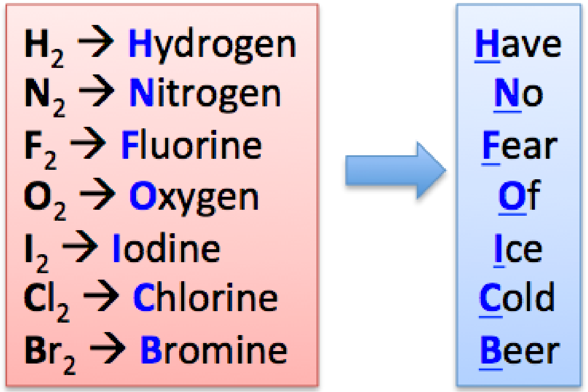
Revisiting Diatomic Elements

Memory aids:

- HIBrONClF
- HOFBrINCl
- **I** **H**ave **N**o **Br**ight **O**r **Cl**ever **F**riends
- **H**ave **N**o **F**ear **Of** **I**ce **C**old **B**eer
- **I** **Br**ing **C**ookies **F**or **O**ur **N**ew **H**ome

…or make your own!

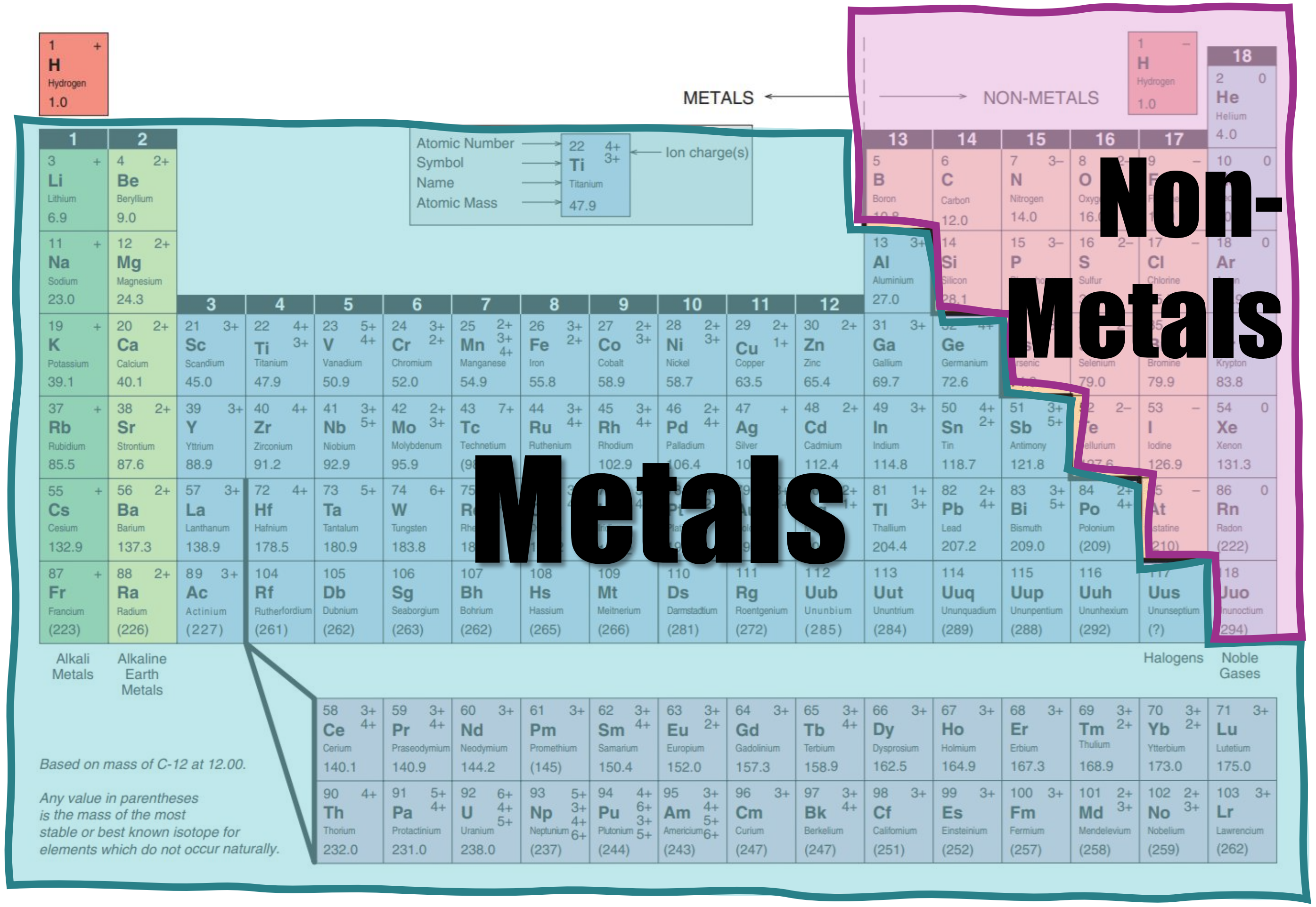




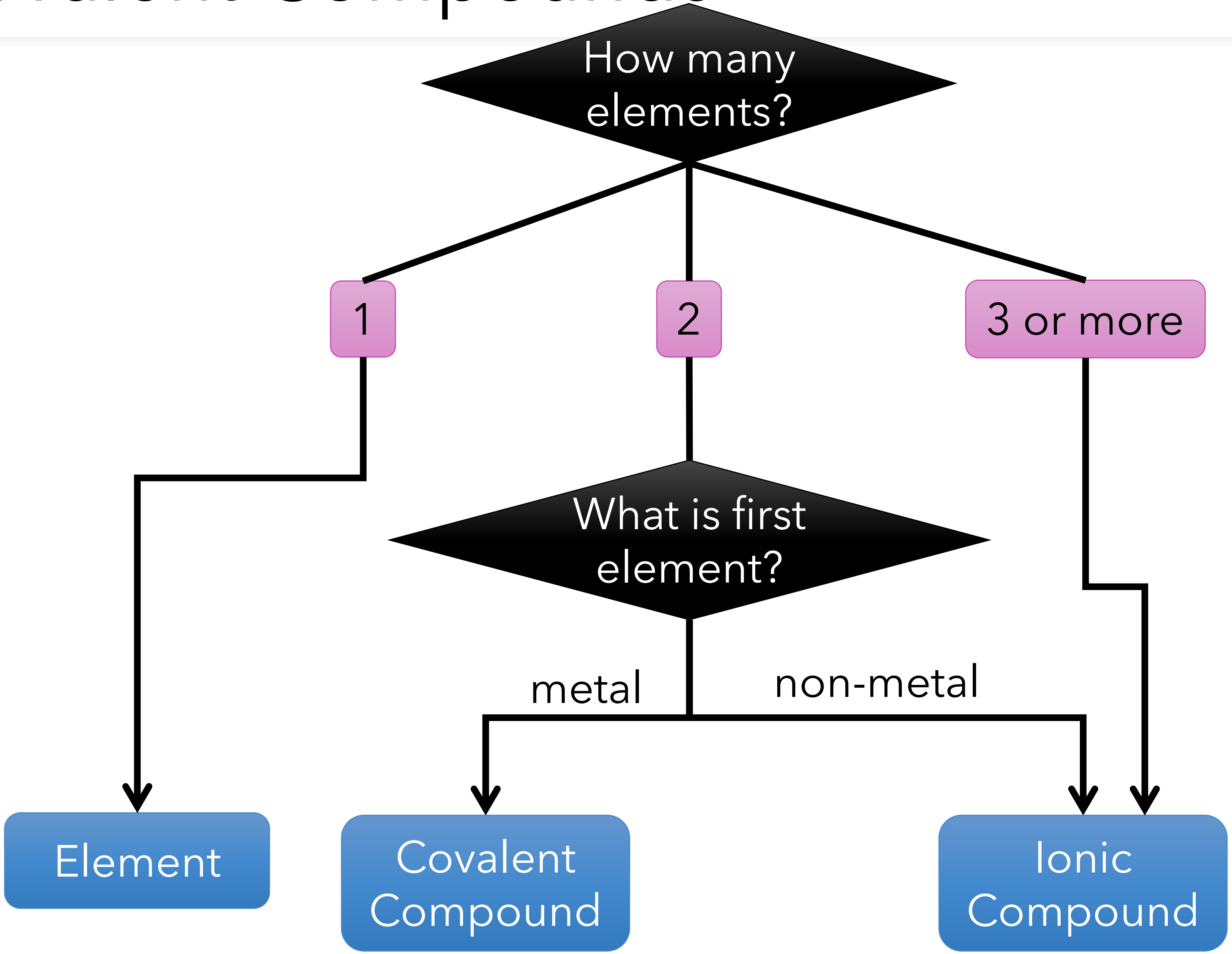
https://study.com/academy/lesson/diatomic-molecule-definition-example.html https://mccord.cm.utexas.edu/chembook/page-nonav.php?chnum=1&sect=12

Identifying Elements, Ionic Compounds, Covalent Compounds

- **Ionic compounds** form when **electrons are transferred** and ions are formed. Usually involves a **metal and a non-metal**.
- **Covalent compounds** form when two (or more) **non-metal** atoms **share electrons**.



Identifying Elements, Ionic Compounds, Covalent Compounds



In Science 9 and 10, you can use the following flowchart to tell apart elements and compounds.

(Note: in nature, many covalent compounds with 3+ elements exist; but we will not learn how to name them.)

Identifying Elements, Ionic Compounds, Covalent Compounds

Chemical	**What is it?**	**Chemical**	**What is it?**
PF_3		NO_2	
$CaCl_2$		Br_2	
Cl_2		NaOH	
TiO		CCl_4	
Al		$MgBr_2$	

Reference

Non-metal Element	"-ide" Ending
N, nitrogen	
O, oxygen	
F, fluorine	
P, phosphorus	
S, sulfur	
Cl, chlorine	

Non-metal Element	"-ide" Ending
Se, selenium	
Br, bromine	
I, iodine	
As, arsenic *	
Te, tellurium *	
At, astatine *	

* uncommon

Arabic Numeral	Roman Numeral	Prefix
1	I	mono
2	II	di
3	III	tri
4	IV	tetra
5	V	penta
6	VI	hexa
7	VII	hepta
8	VIII	octa
9	IX	nona
10	X	deca

Chemical Nomenclature (Naming)

- It is important to have *one* system to name chemical compounds. Why?
  - Scientists can communicate with each other and the public, even in different languages
  - Every compound has a unique name
  - Information/records are accurate and consistent
- IUPAC (International Union of Pure and Applied Chemistry) came up with a naming scheme that is used around the world.

Different Types of Ions

Different Types of Ions

<u>Monovalent ion</u>:

- Can only make one ion (see periodic table)
- Cations: write name of element
- Anions: write name of element with "–ide" ending

Examples:

- Sodium ion = Na^{+}
- Yttrium ion = Y^{3+}
- Bromide ion = Br^{-}
- Oxide ion = O^{2-}

Different Types of Ions

<u>Multivalent Ion</u>:

- An element that can make *<u>multiple possible ions</u>* (see periodic table)
- Metals only
- Must specify charge with Roman numerals

Examples:

- manganese(III) = Mn^{3+}
- manganese(IV) = Mn^{4+}
- copper(I) = Cu^{+}
- vanadium(V) = V^{5+}

Note: manganese and magnesium are *different* elements!

Different Types of Ions

Polyatomic ion:

- Group of non-metal atoms *covalently* bonded with an ionic charge
- Spelling counts!!! (Copy from table)

Examples:

- NH_4^+ = ammonium ion
- PO_4^{3-} = phosphate ion
- PO_3^{3-} = phosphite ion

Polyatomic Ions

Note: Become familiar with these names so you can recognize them quickly in the future.

NAMES, FORMULAE AND CHARGES OF SOME POLYATOMIC IONS

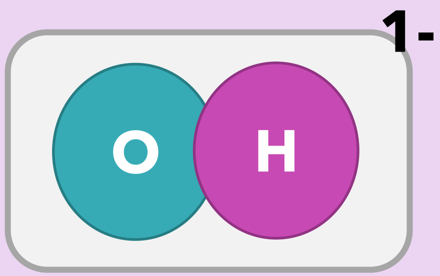
Positive Ions		Negative Ions	
NH_4^+	Ammonium	CH_3COO^-	Acetate
		CO_3^{2-}	Carbonate
		ClO_3^-	Chlorate
		ClO_2^-	Chlorite
		CrO_4^{2-}	Chromate
		CN^-	Cyanide
		$Cr_2O_7^{2-}$	Dichromate
		HCO_3^-	Hydrogen carbonate, bicarbonate
		HSO_4^-	Hydrogen sulfate, bisulfate
		HS^-	Hydrogen sulfide, bisulfide

Positive Ions		Negative Ions	
		HSO_3^-	Hydrogen sulfite, bisulfite
		OH^-	Hydroxide
		ClO^-	Hypochlorite
		NO_3^-	Nitrate
		NO_2^-	Nitrite
		ClO_4^-	Perchlorate
		MnO_4^-	Permanganate
		PO_4^{3-}	Phosphate
		PO_3^{3-}	Phosphite
		SO_4^{2-}	Sulfate
		SO_3^{2-}	Sulfite

Polyatomic Ions

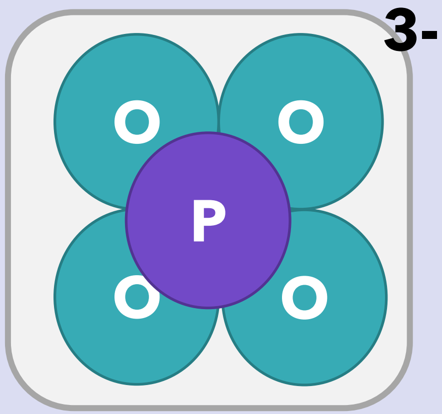
"hydroxide" or "OH^-" is made of an oxygen and hydrogen atom bonded together. Altogether, the structure has a charge of 1-.

e.g. sodium hydroxide: NaOH



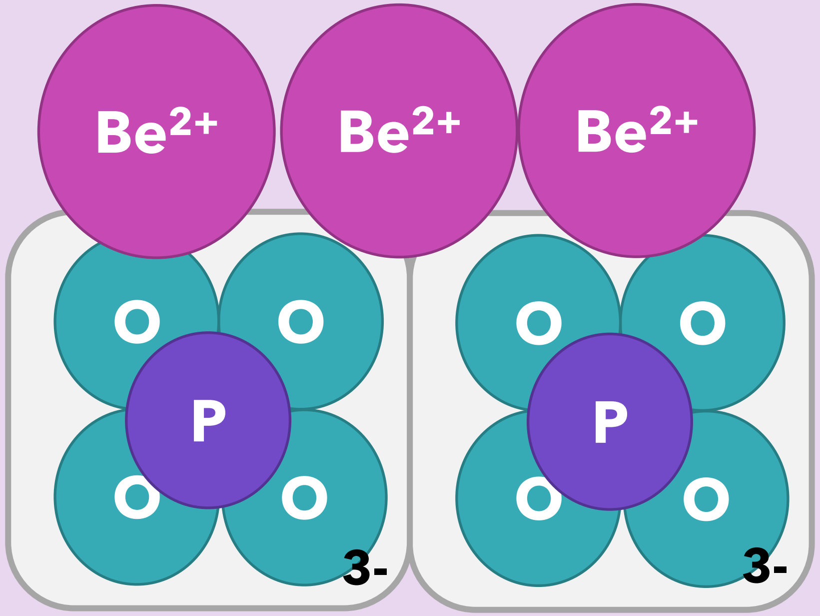
"phosphate" or "PO_4^{3-}" is made of one phosphorus atom and four oxygen atoms bonded together. Altogether, the structure has a charge of 3-.

e.g. sodium phosphate: Na_3PO_4
chromium(II) phosphate: $Cr_3(PO_4)_2$



Polyatomic Ions

To indicate more than one of a polyatomic ion in a compound, use brackets and subscripts.

Chemical Formula	Simplified Model
A subscript outside a bracket applies to the entire polyatomic ion inside the bracket. $Be_3(PO_4)_2$	

Polyatomic Ions

To indicate more than one of a polyatomic ion in a compound, use brackets and subscripts. Treat polyatomic ions as single entities when naming, incl. counting atoms.

Chemical Formula	**Cation**	**Anion**	**Atom Count**
NaOH	Na^{+} x1	OH^{-} x1	Na:1 O:1 H:1
$Mg(OH)_2$	Mg^{2+} x1	OH^{-} x2	Mg:1 O:2 H:1
$Be_3(PO_4)_2$			
$Ti_2(CrO_4)_3$			
$(NH_4)_2Cr_2O_7$			

Naming Ionic Compounds

Intro to Ionic Compound Nomenclature

Cation comes first; anion comes second.

Names of ionic compounds tell you *which ions* are in the compound.

e.g. “sodium chloride” has Na^+ and Cl^- ions.

e.g. “titanium(IV) dichromate” has Ti^{4+} and $Cr_2O_7^{2-}$ ions.

Chemical formulae tell you *how many of each ion* are in the compound, using subscripts.

e.g. “$CaCl_2$” has 1 Ca^{2+} ion and 2 Cl^- ions.

e.g. “$Mn(OH)_2$” has 1 Mn^{4+} ion and 2 OH^- ions.

Intro to Ionic Compound Nomenclature

To write the name or formula of a compound, you must sometimes find out *which ions* are involved, through **charge balancing.**

Rule: The total number of positive charges in an ionic compound must equal the total number of negative charges.

Naming Ionic Compounds

1. Write the ***cation, first***.
2. Write the ***anion with "–ide" ending.***

Chemical Formula	Periodic Table		Name
NaCl	11 + **Na** Sodium 23.0	17 – **Cl** Chlorine 35.5	**sodium chloride**
$MgBr_2$	12 2+ **Mg** Magnesium 24.3	35 – **Br** Bromine 79.9	**magnesium bromide**

Naming Ionic Compounds

1. Write the cation, first.
2. Write the anion with "–ide" ending.

Oh no! Chromium is **multivalent**: it has multiple possible ionic charges. To find out the charge on the chromium ion, we need to do **charge balancing**.

Chemical Formula	Periodic Table	Name
Cr_2O_3	24 Cr Chromium 52.0 (3+, 2+); 8 O Oxygen 16.0 (2–)	???
CrO		???

Naming Ionic Compounds

1. Write the cation, first.

   For metals that can only form one ion (monovalent metals), do not write the ion charge.

   For multivalent metals, determine the ion charge through ***charge balancing***. Then, put the ion charge in ***Roman numerals***, in brackets.

2. Write the anion with "–ide" ending.

Charge Balancing Part 1: Determining Charges of Multivalent Metals

24 $3+$ $2+$
Cr
Chromium
52.0

8 $2-$
O
Oxygen
16.0

Cr_2O_3:	
1) Write out all the ions you have. Leave the charge blank on the multivalent metal.	$Cr^?$ $Cr^?$ O^{2-} O^{2-} O^{2-} We know there are 2 chromium ions and 3 oxygen ions from the subscripts in the formula.
2) *The total number of positive charges in an ionic compound must equal the total number of negative charges.* Determine the charge on the metal ion.	Total: 6 negative charges. Must have 6 positive to balance the charges. Divide by # of chromium ions (2). Therefore, each Cr ion must have a 3+ charge.
3) Write the compound name. Specify the ion charge on the multivalent metal using brackets and Roman numerals.	**chromium(III) oxide**

Charge Balancing Part 1: Determining Charges of Multivalent Metals

24 3+ 2+
Cr
Chromium
52.0

8 2–
O
Oxygen
16.0

CrO:	
1) Write out all the ions you have. Leave the charge blank on the multivalent metal.	$Cr^{?}$ O^{2-} We know there is 1 chromium ion and 1 oxygen ion from the subscripts in the formula.
2) *The total number of positive charges in an ionic compound must equal the total number of negative charges.* Determine the charge on the metal ion.	Total: 2 negative charges. Must have 2 positive to balance the charges. Divide by # of chromium ions (1). Therefore, each Cr ion must have a 2+ charge.
3) Write the compound name. Specify the ion charge on the multivalent metal using brackets and Roman numerals.	**chromium(II) oxide**

Naming Ionic Compounds

1. Write the cation, first.

    For metals that can only form one ion (monovalent metals), do not write the ion charge.

    For multivalent metals, determine the ion charge through charge balancing. Then, put the ion charge in Roman numerals, in brackets.

    If the cation is polyatomic, write it exactly the way it is written in the table.

2. Write the anion with "–ide" ending ***(unless it is polyatomic.)***

Polyatomic Ions

Note: Become familiar with these names so you can recognize them quickly in the future.

NAMES, FORMULAE AND CHARGES OF SOME POLYATOMIC IONS

Positive Ions		Negative Ions	
NH_4^+	Ammonium	CH_3COO^-	Acetate
		CO_3^{2-}	Carbonate
		ClO_3^-	Chlorate
		ClO_2^-	Chlorite
		CrO_4^{2-}	Chromate
		CN^-	Cyanide
		$Cr_2O_7^{2-}$	Dichromate
		HCO_3^-	Hydrogen carbonate, bicarbonate
		HSO_4^-	Hydrogen sulfate, bisulfate
		HS^-	Hydrogen sulfide, bisulfide

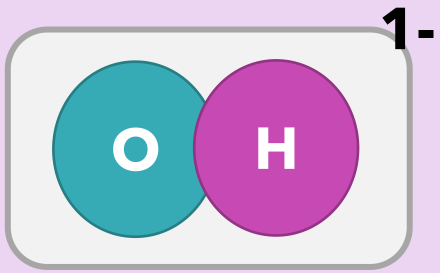
Positive Ions	Negative Ions	
	HSO_3^-	Hydrogen sulfite, bisulfite
	OH^-	Hydroxide
	ClO^-	Hypochlorite
	NO_3^-	Nitrate
	NO_2^-	Nitrite
	ClO_4^-	Perchlorate
	MnO_4^-	Permanganate
	PO_4^{3-}	Phosphate
	PO_3^{3-}	Phosphite
	SO_4^{2-}	Sulfate
	SO_3^{2-}	Sulfite

Polyatomic Ions

Polyatomic ions: ions made of multiple atoms bonded covalently together. They have special names.

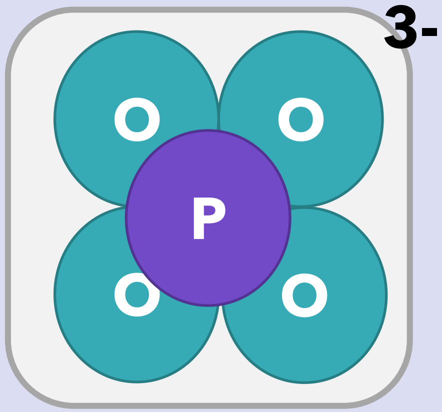
"hydroxide" or "OH^-" is made of an oxygen and hydrogen atom bonded together. Altogether, the structure has a charge of 1-.

e.g. sodium hydroxide: $NaOH$



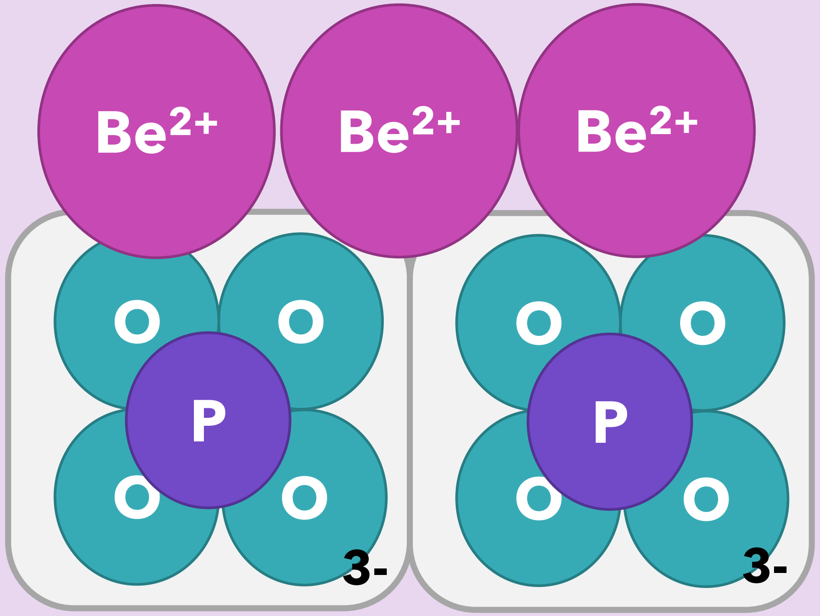
"phosphate" or "PO_4^{3-}" is made of one phosphorus atom and four oxygen atoms bonded together. Altogether, the structure has a charge of 3-.

e.g. sodium phosphate: Na_3PO_4
chromium(II) phosphate: $Cr_3(PO_4)_2$



Polyatomic Ions

To indicate more than one of a polyatomic ion in a compound, use brackets and subscripts.

Chemical Formula	**Simplified Model**
A subscript outside a bracket applies to the entire polyatomic ion inside the bracket. $Be_3(PO_4)_2$	

Polyatomic Ions

To indicate more than one of a polyatomic ion in a compound, use brackets and subscripts. Treat polyatomic ions as single entities when naming, incl. counting atoms.

Chemical Formula	Cation	Anion	Atom Count
NaOH	Na^{+}	OH^{-}	Na:1 O:1 H:1
$Mg(OH)_2$	Mg^{2+}	OH^{-} x2	Mg:1 O:2 H:1
$Be_3(PO_4)_2$	Be^{2+} x3	PO_4^{2-} x2	Be:3 P:2 O:8
$Ti_2(CrO_4)_3$	Ti^{3+} x2	CrO_4^{2-} x3	Ti:2 Cr:3 O:12

Rules for Naming Ionic Compounds (FINAL)

1. Write the ***cation, first***.

   For metals that can only form one ion (monovalent metals), do not write the ion charge.

   For multivalent metals, determine the ion charge through ***charge balancing***. Then, put the ion charge in ***Roman numerals***, in brackets.

   If the cation is polyatomic, write it exactly the way it is written in the table.

2. Write the anion with ***"–ide" ending*** (unless it is polyatomic.)

Naming with Polyatomic Ions: Examples

Chemical Formula	Periodic Table	Name
$Mg(OH)_2$	12 2+ **Mg** Magnesium 24.3 HSO_3^- Hydrogen sulfite, bisulfite OH^- Hydroxide ClO^- Hypochlorite	**magnesium hydroxide**
$(NH_4)_2S$	**Positive Ions** NH_4^+ Ammonium 16 2– **S** Sulfur 32.1	**ammonium sulfide**

Naming with Polyatomic Ions: Examples

Chemical Formula	Periodic Table	Name
$Sc(HSO_3)_3$	21 3+ **Sc** Scandium 45.0 HSO_4^- Hydrogen sulfate, bisulfate HS^- Hydrogen sulfide, bisulfide HSO_3^- Hydrogen sulfite, bisulfite	**1. scandium hydrogen sulfite** *OR* **2. scandium bisulfite** ~~scandium hydrogen sulfite, bisulfite~~

Naming with Polyatomic Ions: Examples

22 4+
Ti 3+
Titanium
47.9

ClO_2^-	Chlorite
CrO_4^{2-}	Chromate
CN^-	Cyanide

$Ti_2(CrO_4)_3$:	
1) Write out all the ions you have. Leave the charge blank on the multivalent metal.	$Ti^?$ $Ti^?$ CrO_4^{2-} CrO_4^{2-} CrO_4^{2-}
2) *The total number of positive charges in an ionic compound must equal the total number of negative charges.* Determine the charge on the metal ion.	Total: 6 negative charges. Must have 6 positive to balance the charges. Divide by # of titanium ions (2). Therefore, each Ti ion must have a 3+ charge.
3) Write the compound name. Specify the ion charge on the multivalent metal using brackets and Roman numerals. Spell the polyatomic ion exactly as it is spelled in the reference sheet.	**titanium(III) chromate**

Writing Formulas of Ionic Compounds

Intro to Ionic Compound Nomenclature

Names of ionic compounds tell you *which ions* are in the compound. The cation comes first; the anion comes second.

To write a chemical formula of an ionic compound, you must find out how many of each ion is involved, through **charge balancing.**

Rule: The total number of positive charges in an ionic compound must equal the total number of negative charges.

Writing Formulas of Ionic Compounds (v1)

1. Write down each ion with its charge.
2. Add more of the ions to balance the charges: the total number of positive and negative charges must be equal.
3. Write your formula with subscripts.

   To indicate more than one of a polyatomic ion, use brackets with the subscript outside.

Writing Chemical Formulas: Examples (v1)

20 2+
Ca
Calcium
40.1

15 3–
P
Phosphorus
31.0

<table>
<tr><th colspan="2">calcium phosphide</th></tr>
<tr><td>1) Write down each ion with its charge.</td><td rowspan="2">Ca^{2+} P^{3-}
Ca^{2+} P^{3-}
Ca^{2+}</td></tr>
<tr><td>2) Add more of the ions to balance the charges: the total number of positive and negative charges must be equal.</td></tr>
<tr><td>3) Write your formula with subscripts.</td><td>$\mathbf{Ca_3P_2}$</td></tr>
</table>

Writing Chemical Formulas: Examples (v1)

24	$3+$
Cr	$2+$
Chromium	
52.0	

HSO_3^-	Hydrogen sulf
OH^-	Hydroxide
ClO^-	Hypochlorite

chromium(II) hydroxide	
1) Write down each ion with its charge.	Cr^{2+} OH^-
2) Add more of the ions to balance the charges: the total number of positive and negative charges must be equal.	OH^-
3) Write your formula with subscripts.	$\mathbf{Cr(OH)_2}$

Writing Formulas of Ionic Compounds (v2)

1. Write down each ion with its charge.
2. Write the chemical formula by writing the cation first and the anion second. Then, “criss-cross” the charges to become the subscripts.
3. Reduce the subscripts if both divisible by the same number.

Writing Chemical Formulas: Examples (v2)

20 2+
Ca
Calcium
40.1

15 3–
P
Phosphorus
31.0

calcium phosphide	
1) Write down each ion with its charge.	Ca^{2+} P^{3-}
2) Write the chemical formula by writing the cation first and the anion second. Then, "criss-cross" the charges to become the subscripts.	Ca_3P_2
3) Reduce the subscripts if both divisible by the same number.	2 and 3 do not have a common factor. Therefore, $\mathbf{Ca_3P_2}$ is our final answer.

Writing Chemical Formulas: Examples (v2)

24 3+ 2+
Cr
Chromium
52.0

HSO_3^-	Hydrogen sulf
OH^-	Hydroxide
ClO^-	Hypochlorite

chromium(II) hydroxide	
1) Write down each ion with its charge.	Cr^{2+} OH^-
2) Write the chemical formula by writing the cation first and the anion second. Then, "criss-cross" the charges to become the subscripts.	$Cr(OH)_2$
3) Reduce the subscripts if both divisible by the same number.	1 and 2 do not have a common factor. Therefore, **$Cr(OH)_2$** is our final answer.

Writing Chemical Formulas: Examples (v2)

12	2+
Mg	
Magnesium	
24.3	

CH_3COO^-	Acetate
CO_3^{2-}	Carbonate
ClO_3^-	Chlorate

magnesium carbonate	
1) Write down each ion with its charge.	Mg^{2+} CO_3^{2-}
2) Write the chemical formula by writing the cation first and the anion second. Then, "criss-cross" the charges to become the subscripts.	$Mg_2(CO_3)_2$
3) Reduce the subscripts if both divisible by the same number.	2 and 2 are both divisible by 2. Rewrite formula as **$MgCO_3$**.

Writing Chemical Formulas: Examples (v2)

25 | 2+ 3+ 4+
Mn
Manganese
54.9

PO_3^{3-}	Phosphite
SO_4^{2-}	Sulfate
SO_3^{2-}	Sulfite

manganese(IV) sulfate	
1) Write down each ion with its charge.	Mn^{4+} SO_4^{2-} $Mn_2(SO_4)_4$
2) Write the chemical formula by writing the cation first and the anion second. Then, "criss-cross" the charges to become the subscripts.	
3) Reduce the subscripts if both divisible by the same number.	4 and 2 are both divisible by 2. Rewrite formula as $\mathbf{Mn(SO_4)_2}$.

Naming and Writing Formulas: Covalent Compounds

Naming Binary Covalent Compounds

- Binary covalent compound: a covalent compound containing only two element
- Names and formulas of covalent compounds *both* tell you:
  - Which elements
  - How many atoms of each element

Naming Binary Covalent Compounds

1. Write the first element.
2. Write the second element with "–ide" ending.
3. Add **prefixes** to show how many of each element there is.
   - *Do not add "mono-" to first element.*
   - *If adding "mono-" to "-oxide", write "monoxide" instead.*

e.g. O_2F_2 **di oxygen difluoride**

e.g. PF_3 **phosphorus tri fluoride**

e.g. N_2O **di nitrogen monoxide**

Note: All compound names (covalent *and* ionic) are lowercase.

Naming Binary Covalent Compounds

Covalent compounds with special names (must memorize):

NH_4 = ammonia

H_2O = water

CH_4 = methane

NH_4^+ (ammonium ion) and NH_4 (ammonia) are ***not the same!!!***

Chemical Formulas of Binary Covalent Compounds

1. Identify the elements involved. Write their symbols.
2. Use the prefixes to determine the number of each element in the compound. Write as subscripts.

e.g. tetraphosphorus pentaoxide

P_4O_5

e.g. nitrogen triiodide

$N\,I_3$

e.g. xenon hexafluoride

$Xe\,F_6$

More Practice: Binary Covalent Compounds

Chemical Formula	**Compound Name**
CO_2	
CO	
CCl_4	
P_4O_5	
	diphosphorus pentaoxide
	xenon hexafluoride

Fruit Tart Case Study

You are making fruit tarts for a party. Unfortunately, after you are finished, you see an Instagram picture that makes you want to rearrange your fruit tarts. You need 3 finished raspberry/blackberry tarts in total. How many of each tart will you start with? What will you be left with?

6 raspberries each

+

1 blackberry each

→

2 raspberries +
1 blackberry each

https://www.mypetitejoys.com/no-bake-blackberry-mini-tarts/ https://leaf.nutrisystem.com/recipes/skinny-mini-fruit-tart/ https://www.youtube.com/watch?v=mfGNSkXOUHo

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6 raspberries each

+

1 blackberry each

→

2 raspberries +
1 blackberry each

+

fruitless tart

Discuss: approaches and strategies in completing this problem

https://veenaazmanov.com/homemade-mini-tart-shells/

https://www.mypetitejoys.com/no-bake-blackberry-mini-tarts/ https://leaf.nutrisystem.com/recipes/skinny-mini-fruit-tart/ https://www.youtube.com/watch?v=mfGNSkXOUHo

Fruit Tart Case Study

6 raspberries each

+

1 blackberry each

→

2 raspberries +
1 blackberry each

+

fruitless tart

$$\underline{\ 1\ }\,Rb_6T + \underline{\ 3\ }\,BbT \rightarrow \underline{\ 3\ }\,Rb_2BbT + \underline{\ 1\ }\,T$$

Legend

Rb = “raspberry” element
Bb = “blackberry” element
T = “tart” element

Follow-up: Now, suppose that you need 12 tarts instead of 3. How many raspberry and blackberry tarts do you start with?

https://www.mypetitejoys.com/no-bake-blackberry-mini-tarts/
https://leaf.nutrisystem.com/recipes/skinny-mini-fruit-tart/
https://veenaazmanov.com/homemade-mini-tart-shells/
https://www.youtube.com/watch?v=mfGNSkXOUHo

Balancing Chemical Equations

Why balance?

- Chemical "recipes": how much do you put in? how much do you expect to yield?
- **Conservation of mass**: no atoms are ever created or destroyed

https://www.budgetbytes.com/

Balancing Chemical Equations: Vocabulary

Balancing chemical formulas involves adding **coefficients** in front of elements and compounds until ***the total atoms in the reactants equals the products***.

coefficients (balancing numbers)

$$2H_2 + O_2 \rightarrow 2H_2O$$

subscripts (don't mess with these)

Balancing Chemical Equations: Vocabulary

Balancing chemical formulas involves adding **coefficients** in front of elements and compounds until ***the total atoms in the reactants equals the products***.

- **Element:** made of one type of atom
- **Compound:** made of two or more types of atoms

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Balancing Chemical Equations: Vocabulary

Balancing chemical formulas involves adding **coefficients** in front of elements and compounds until ***the total number of atoms of each element in the reactants equals the products***.

Reactants: what goes into the reaction

Products: what comes out of the reaction

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Balancing Chemical Equations: Tips

- **Goal: the number of atoms of each element in the reactants equals the products.** Guess and check until this happens!
- Remember your diatomic elements: **H, I, Br, O, N, Cl, F**
- Balance atoms in compounds first. Save elements for last.
- If the same polyatomic ion appears in the reactants *and* products, you can often treat it as a group of atoms instead of splitting it up.
- At the end, reduce all coefficients to lowest whole-number terms.

Note: balancing can be frustrating at first. Practice, practice, practice!

Balancing Examples (easy)

Note: Do not write a coefficient if there is only "1" of that element or compound.

1. $\_\_\_\_ N_2 + \underline{\ 3\ } H_2 \rightarrow \underline{\ 2\ } NH_3$

2. $\underline{\ 2\ } NaCl + \_\_\_\_ F_2 \rightarrow \underline{\ 2\ } NaF + \_\_\_\_ Cl_2$

3. $\underline{\ 4\ } P + \underline{\ 5\ } O_2 \rightarrow \underline{\ 2\ } P_2O_5$

4. $\underline{\ 2\ } Ag_2O \rightarrow \underline{\ 4\ } Ag + \_\_\_\_ O_2$

Balancing Examples (medium)

Treat polyatomic ions as groups if they appear in reactants and products (e.g. #2 & #3 but not #5)

5. $\underline{\;2\;}$ NaBr + ___ $CaF_2 \rightarrow$ $\underline{\;2\;}$ NaF + ___ $CaBr_2$

6. ___ $FeCl_3$ + $\underline{\;3\;}$ NaOH $\rightarrow$ ___ $Fe(OH)_3$ + $\underline{\;3\;}$ NaCl

7. ___ H_2SO_4 + $\underline{\;2\;}$ $NaNO_2 \rightarrow$ $\underline{\;2\;}$ HNO_2 + ___ Na_2SO_4

8. $\underline{\;6\;}$ CO_2 + $\underline{\;6\;}$ $H_2O \rightarrow$ ___ $C_6H_{12}O_6$ + $\underline{\;6\;}$ O_2

9. $\underline{\;2\;}$ HCl + ___ $CaCO_3 \rightarrow$ ___ $CaCl_2$ + ___ H_2O + ___ CO_2

Balancing Examples (hard)

10. ___ C_3H_8 + 5 O_2 → 3 CO_2 + 4 H_2O

11. 2 C_6H_{14} + 19 O_2 → 12 CO_2 + 14 H_2O

12. 2 C_8H_{18} + 25 O_2 → 16 CO_2 + 18 H_2O

Make sure to balance the element (O_2) last!

Trick for Combustion Reactions (e.g. #10-12)

1. **Balance every atom except oxygen.**

$$\_\_\_ C_6H_{14} + \_\_\_ O_2 \rightarrow \underline{\ 6\ } CO_2 + \underline{\ 7\ } H_2O$$

2. **Find out how many oxygen atoms you need the __O_2 to contribute. Divide that number by 2. This is your *temporary* coefficient for O_2.**

$$\_\_\_ C_6H_{14} + \underline{\tfrac{19}{2}} O_2 \rightarrow \underline{\ 6\ } CO_2 + \underline{\ 7\ } H_2O$$

$6CO_2$ has 12 oxygen atoms.
$7H_2O$ has 7 oxygen atoms.
In total, there are 19 oxygen atoms in the products.

3. **You are not allowed to have fractional coefficients in your final answer. Multiply all the coefficients by 2.**

$$\underline{\ 2\ } C_6H_{14} + \underline{\ 19\ } O_2 \rightarrow \underline{\ 12\ } CO_2 + \underline{\ 14\ } H_2O$$

Resources

- Naming and Writing Chemical Formulas
  - Mr. Carman's Blog (generates quizzes) https://www.kentschools.net/ccarman/cp-chemistry/practice-quizzes/compound-naming/
  - Mr. Eisley (list of other resources to practice http://www.mreisley.com/nomenclature-practice.html
  - ChemFiesta (worksheets with answers) https://chemfiesta.org/2015/01/13/naming-worksheets/
- Balancing Chemical Equations
  - TemplateLAB (explanations and many worksheets with answers) https://templatelab.com/balancing-equations-worksheet/

Practice

Classify as ionic or covalent. Then, name the following compounds:

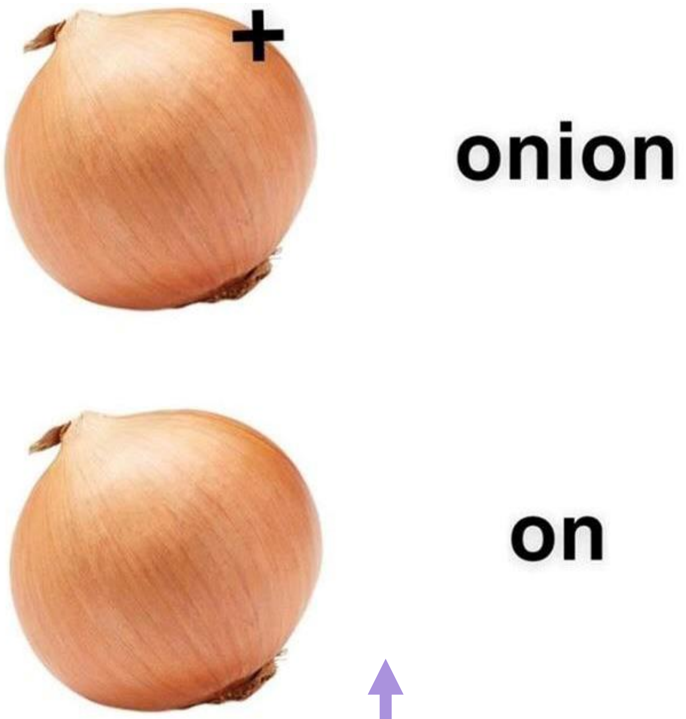
Formula	Name
CO_2	
Na_2O	
CrF_3	
N_2Br_3	
MnO_2	

Try to classify as ionic or covalent. How are these compounds different from what we have seen so far?

Formula	Name
$MgCO_3$	magnesium carbonate
$Ca(CH_3COO)_2$	calcium acetate
NH_4Br	ammonium bromide
KCN	potassium cyanide

Whenever you have **more than 2 elements** in your ionic compound, you are dealing with a polyatomic ion. Polyatomic ions are almost always anions (except NH_4^+).

Ionic Compound Formation



(This is a joke. Not to be confused with mnemonic on previous page.)

https://www.reddit.com/r/chemistrymemes/comments/ccsxov/high_iq_meme/

https://www.chemistryjokes.com/jokes/cation/